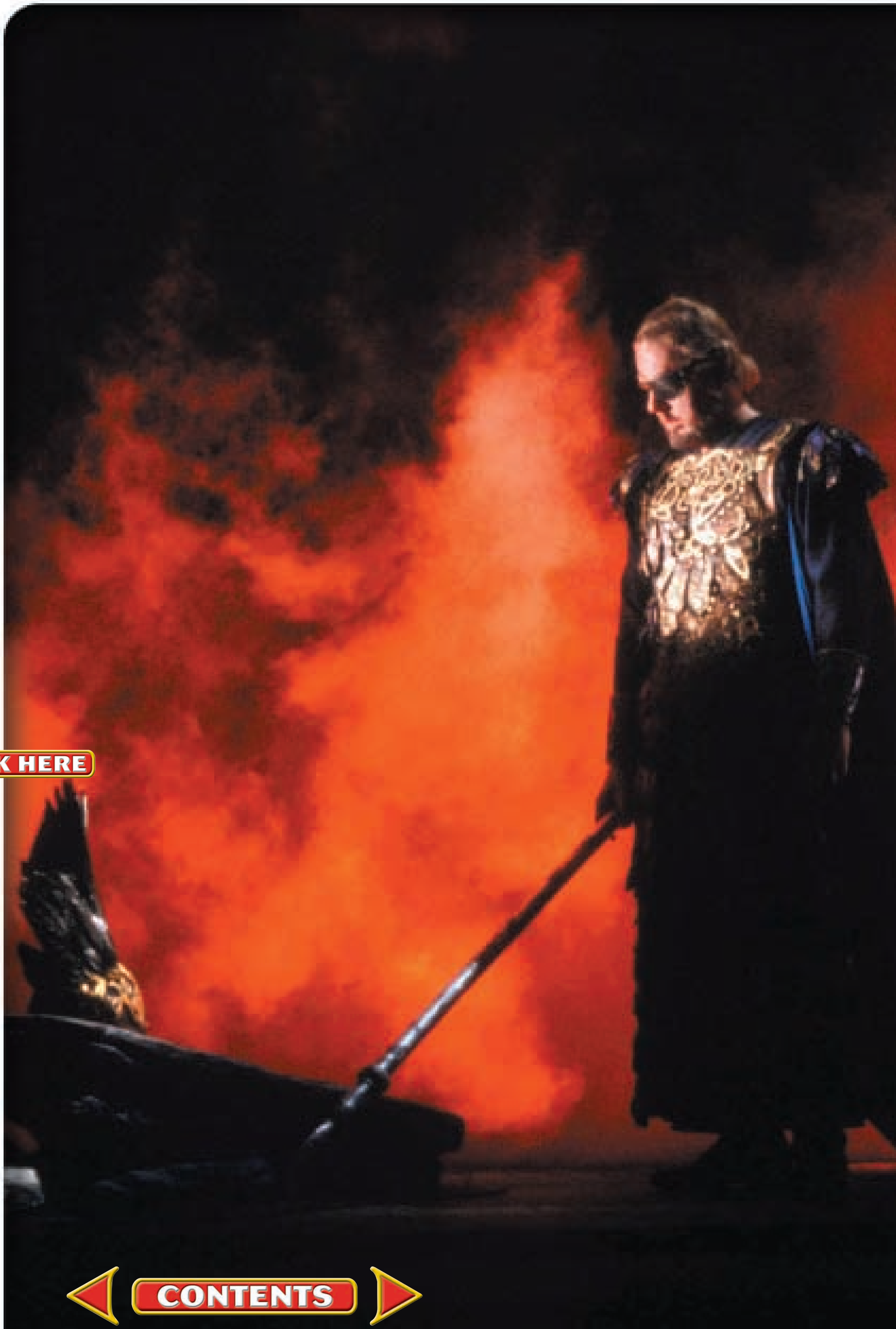
CHAPTER 13

States of Matter



What You'll Learn

- You will use the kinetic-molecular theory to explain the physical properties of gases, liquids, and solids.
- You will compare types of intermolecular forces.
- You will explain how kinetic energy and intermolecular forces combine to determine the state of a substance.
- You will describe the role of energy in phase changes.

Why It's Important

Water collects on a bathroom mirror as you shower, a full bottle of water shatters in a freezer, and a glass object breaks when it is dropped. You will be able to explain such familiar events after you learn more about the different states of matter.

CLICK HERE

Visit the Chemistry Web site at **science.glencoe.com** to find links about the states of matter and phase changes.

Solid carbon dioxide is called dry ice. At room temperature, dry ice is used to create the illusion of fog on stage.

DISCOVERY LAB

Materials

pin
600-mL beaker
400 mL water
detergent
dropper

Defying Density

You know that an object sinks or floats in water based on its density. In this activity, you will explore an exception to this rule.

Safety Precautions

Be careful handling the pin, which has a sharp point.

Procedure

1. Pour about 400 mL of water into a 600-mL beaker. Float the pin on the surface of the water.
2. Use a dropper to add one drop of water containing detergent to the beaker. Place the drop on the water surface near the wall of the beaker. Observe what happens.

Analysis

Is a metal pin likely to be more or less dense than water? How does the shape of the pin help it to float? Hypothesize about the reason for the pin's behavior before and after you added the detergent.

Section 13.1

Gases

Objectives

- **Use** the kinetic-molecular theory to explain the behavior of gases.
- **Describe** how mass affects the rates of diffusion and effusion.
- **Explain** how gas pressure is measured and **calculate** the partial pressure of a gas.

Vocabulary

kinetic-molecular theory
elastic collision
temperature
diffusion
Graham's law of effusion
pressure
barometer
pascal
atmosphere
Dalton's law of partial pressures

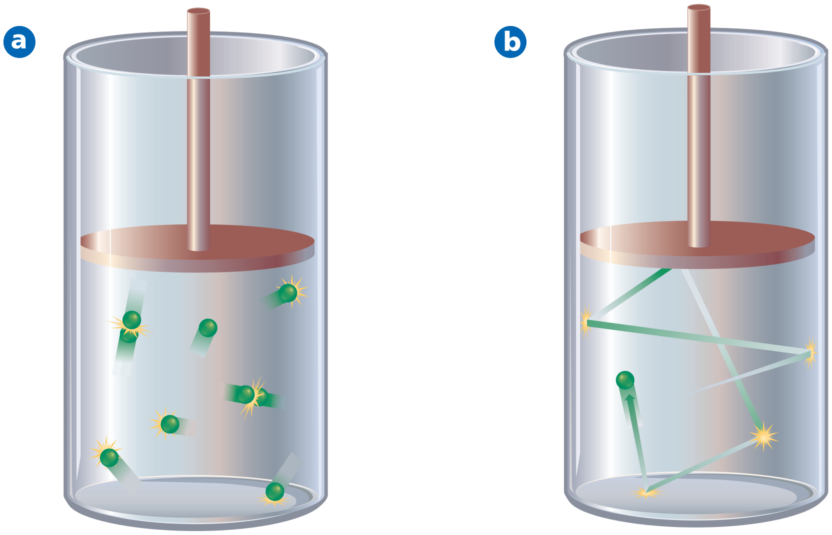
You have learned that the types of atoms present (composition) and their arrangement (structure) determine the chemical properties of matter. Composition and structure also affect the physical properties of liquids and solids. Based solely on physical appearance, you can distinguish water from mercury or gold from graphite. By contrast, substances that are gases at room temperature usually display similar physical properties despite their different compositions. Why is there so little variation in behavior among gases? Why are the physical properties of gases different from those of liquids and solids?

The Kinetic-Molecular Theory

Flemish physician Jan Baptista Van Helmont (1577–1644) used the Greek word *chaos*, which means without order, to describe those products of reactions that had no fixed shape or volume. From the word *chaos* came the term *gas*. By the eighteenth century, scientists knew how to collect gaseous products by displacing water. Now they could observe and measure properties of individual gases. About 1860, Ludwig Boltzmann and James Maxwell, who were working in different countries, each proposed a model to explain the properties of gases. That model is the kinetic-molecular theory. Because all of the gases known to Boltzmann and Maxwell contained molecules, the name of the model refers to molecules. The word *kinetic* comes from a Greek word meaning "to move." Objects in motion have energy called kinetic energy. The **kinetic-molecular theory** describes the behavior of gases in terms of particles in motion. The model makes several assumptions about the size, motion, and energy of gas particles.

Figure 13-1

a Based on its size, is a gas particle more likely to collide with another particle or with the walls of its container?
b What happens to the path of a gas particle after a collision?



Particle size Gases consist of small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space. Because gas particles are far apart, there are no significant attractive or repulsive forces among them.

Particle motion Gas particles are in constant, random motion. Particles move in a straight line until they collide with other particles or with the walls of their container, as shown in **Figure 13-1.** Collisions between gas particles are elastic. An **elastic collision** is one in which no kinetic energy is lost. Kinetic energy may be transferred between colliding particles, but the total kinetic energy of the two particles does not change.

Particle energy Two factors determine the kinetic energy of a particle: mass and velocity. The kinetic energy of a particle can be represented by the equation

$$KE = \frac{1}{2}mv^2$$

in which KE is kinetic energy, m is the mass of the particle, and v is its velocity. Velocity reflects both the speed and the direction of motion. In a sample of a single gas, all particles have the same mass but all particles do not have the same velocity. Therefore, all particles do not have the same kinetic energy. Kinetic energy and temperature are related. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter. At a given temperature, all gases have the same average kinetic energy.

Figure 13-2

The air in a life jacket allows the person wearing the jacket to float on the water.

Explaining the Behavior of Gases

Kinetic-molecular theory can help explain the behavior of gases. For example, the constant motion of gas particles allows a gas to expand until it fills its container, such as the flotation device in **Figure 13-2.** What property of gases makes it possible for an air-filled flotation device to work?

Low density Remember that density is mass per unit volume. The density of chlorine gas is 2.95×10^{-3} g/mL at 20°C; the density of solid gold is 19.3 g/mL. Gold is more than 6500 times as dense as chlorine. This large difference cannot be due only to the difference in mass between gold atoms and chlorine molecules

(about 3:1). As the kinetic-molecular theory states, a great deal of space exists between gas particles. Thus, there are fewer chlorine molecules than gold atoms in the same volume.

Compression and expansion If you squeeze a pillow made of foam, you can compress it; that is, you can reduce its volume. The foam contains air pockets. The large amount of empty space between the air particles in those pockets allows the air to be easily pushed into a smaller volume. When you stop squeezing, the random motion of air particles fills the available space and the pillow expands to its original shape. **Figure 13-3** illustrates what happens to the density of a gas in a container as it is compressed and as it is allowed to expand.

Diffusion and effusion According to the kinetic-molecular theory, there are no significant forces of attraction between gas particles. Thus, gas particles can flow easily past each other. Often, the space into which a gas flows is already occupied by another gas. The random motion of the gas particles causes the gases to mix until they are evenly distributed. **Diffusion** is the term used to describe the movement of one material through another. The term may be new, but you are probably familiar with the process. If you are in the den, can you tell when someone sprays perfume in the bedroom? Perfume particles released in the bedroom diffuse through the air until they reach the den. Particles diffuse from an area of high concentration (the bedroom) to one of low concentration (the den).

The rate of diffusion depends mainly on the mass of the particles involved. Lighter particles diffuse more rapidly than heavier particles. Recall that different gases at the same temperature have the same average kinetic energy as described by the equation $KE = 1/2mv^2$. However, the mass of gas particles varies from gas to gas. For lighter particles to have the same average kinetic energy as heavier particles, they must, on average, have a greater velocity.

Effusion is a process related to diffusion. During effusion, a gas escapes through a tiny opening. What happens when you puncture a container such as a balloon or a tire? In 1846, Thomas Graham did experiments to measure the rates of effusion for different gases at the same temperature. Graham designed his experiment so that the gases effused into a vacuum—a space containing no matter. He discovered an inverse relationship between effusion rates and molar mass. **Graham's law of effusion** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

Graham's law also applies to rates of diffusion, which is logical because heavier particles diffuse more slowly than lighter particles at the same temperature. Using Graham's law, you can set up a proportion to compare the diffusion rates for two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

For example, consider the gases ammonia (NH_3) and hydrogen chloride (HCl). Which gas diffuses faster? Example Problem 13-1 and the photograph on the next page provide quantitative and visual comparisons of the diffusion rates for ammonia and hydrogen chloride.

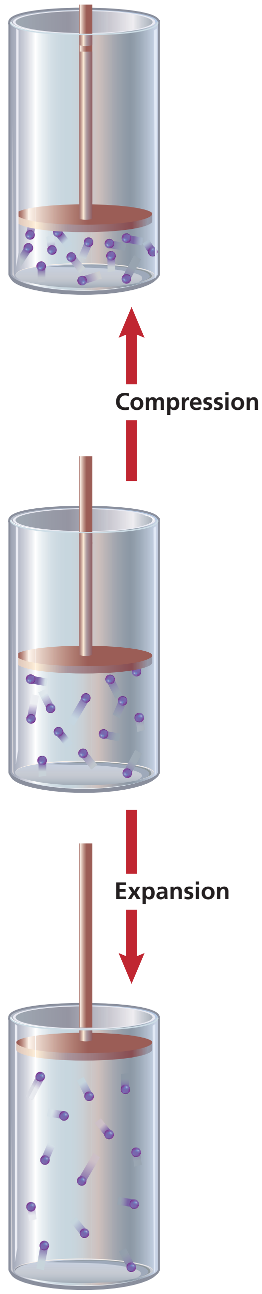


Figure 13-3

The volume of the top cylinder is half that of the middle cylinder. The volume of the bottom cylinder is twice that of the middle cylinder. Compare the density of the gas in the top cylinder with its density in the bottom cylinder.

EXAMPLE PROBLEM 13-1

Finding a Ratio of Diffusion Rates

Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

When molecules of gaseous HCl (from a bottle of hydrochloric acid) and NH_3 (from a bottle of aqueous ammonia) meet, they react to form the white solid ammonium chloride, NH_4Cl.

1. Analyze the Problem

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known	Unknown
$\text{molar mass}_{HCl} = 36.5\text{ g/mol}$	ratio of diffusion rates = ?
$\text{molar mass}_{NH_3} = 17.0\text{ g/mol}$	

2. Solve for the Unknown

Substitute the known values into Graham's equation and solve.

$$\frac{\text{Rate}_{NH_3}}{\text{Rate}_{HCl}} = \sqrt{\frac{\text{molar mass}_{HCL}}{\text{molar mass}_{NH_3}}}$$

$$= \sqrt{\frac{36.5\ \cancel{\text{g/mol}}}{17.0\ \cancel{\text{g/mol}}}} = \sqrt{2.15} = 1.47$$

The ratio of diffusion rates is 1.47.

3. Evaluate the Answer

Ammonia molecules diffuse about 1.5 times as fast as hydrogen chloride molecules. This ratio is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer does, too.

PRACTICE PROBLEMS

1. Calculate the ratio of effusion rates for nitrogen (N_2) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide (CO) and carbon dioxide (CO_2).
3. What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

Practice!

For more practice with diffusion problems, go to **Supplemental Practice Problems** in Appendix A.

Gas Pressure

If you were to walk across deep snow in boots, you would probably sink into the snow with each step. With a pair of snowshoes like those in **Figure 13-4,** you would be less likely to sink. In each case, the force with which you press down on the snow is related to your mass. With snowshoes, the force would be spread out over a larger area. Therefore, the pressure on any given area of snow would be reduced. **Pressure** is defined as force per unit area. How do the size and style of shoes you wear affect the force you exert on a surface?

Gas particles exert pressure when they collide with the walls of their container. Because an individual gas particle has little mass, it can exert little pressure. However, there are about 10^{22} gas particles in a liter container. With this many particles colliding, the pressure can be substantial. In Chapter 14, you will learn how temperature, volume, and number of moles affect the pressure that a gas exerts.

CONTENTS

Earth is surrounded by an atmosphere that extends into space for hundreds of kilometers. Because the particles in air move in every direction, they exert pressure in all directions. This pressure is called atmospheric pressure, or air pressure. At the surface of Earth, air pressure is approximately equal to the pressure exerted by a 1-kilogram mass on a square centimeter. Air pressure varies at different points on Earth. At the top of a mountain, the mass of the air column pressing down on a square centimeter of Earth is less than the mass of the air column at sea level. Thus, the air pressure at higher altitudes is slightly lower than at sea level.

Measuring air pressure Italian physicist Evangelista Torricelli (1608–1647) was the first to demonstrate that air exerted pressure. He had noticed that water pumps were unable to pump water higher than about ten meters. He hypothesized that the height of a column of liquid would vary with the density of the liquid. To test this idea, Torricelli designed the equipment shown in **Figure 13-5a.** He filled a thin glass tube that was closed at one end with mercury. While covering the open end so that air could not enter, he inverted the tube and placed it (open end down) in a dish of mercury. The open end was below the surface of the mercury in the dish. The height of the mercury in the tube fell to about 75 cm. Mercury is about 13.6 times denser than water. Did the results of his experiment support Torricelli's hypothesis?

The device that Torricelli invented is called a barometer. A **barometer** is an instrument used to measure atmospheric pressure. As Torricelli showed, the height of the mercury in a barometer is always about 760 mm. The exact height of the mercury is determined by two forces. Gravity exerts a constant downward force on the mercury. This force is opposed by an upward force exerted by air pressing down on the surface of the mercury. Changes in air temperature or humidity cause air pressure to vary. An increase in air pressure causes the mercury to rise; a decrease causes the mercury to fall.

A manometer is an instrument used to measure gas pressure in a closed container. In a manometer, a flask is connected to a U-tube that contains mercury. In **Figure 13-5b,** there is no gas in the flask. The mercury is at the same height in each arm of the U-tube. In **Figure 13-5c**, there is gas in the flask. When the valve between the flask and the U-tube is opened, gas particles diffuse out of the flask into the U-tube. The released gas particles push down on the mercury in the tube. What happens to the height of the mercury in each arm of the U-tube? The difference in the height of the mercury in the two arms is used to calculate the pressure of the gas in the flask.

Figure 13-4

Snowshoes reduce the force on a given area. What devices other than snowshoes can people use to travel across deep snow?

Figure 13-5

a A barometer measures air pressure. **b** A manometer measures the pressure of an enclosed gas. Before gas is released into the U-tube, the mercury is at the same height in each arm. **c** After gas is released into the U-tube, the heights in the two arms are no longer equal.

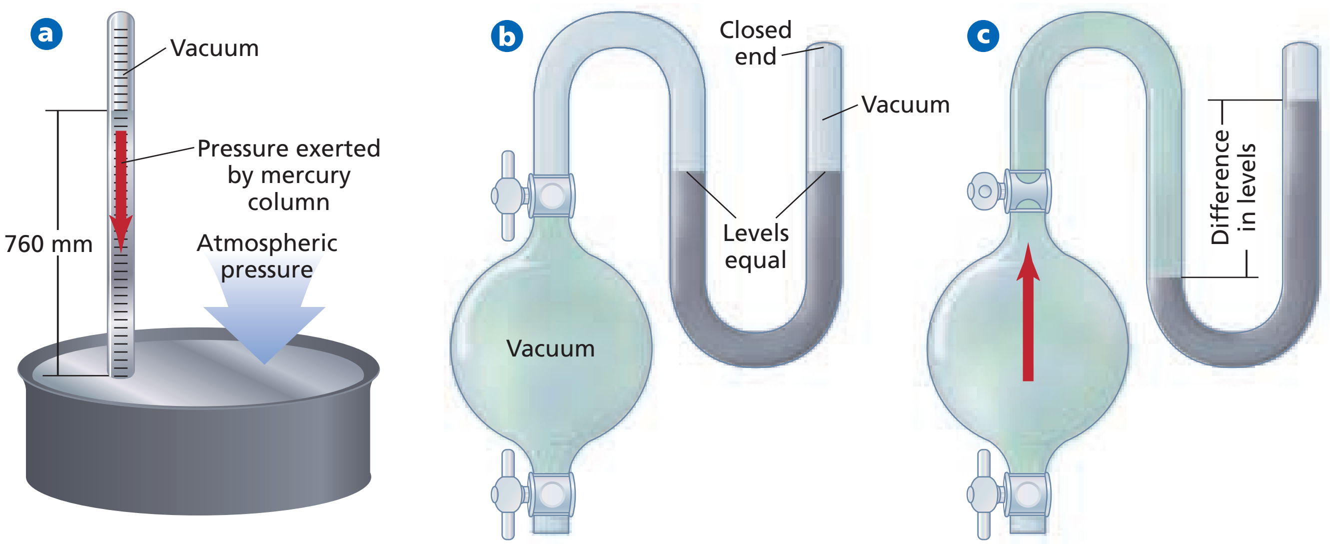


Table 13-1

Comparison of Pressure Units		
Unit	**Compared with 1 atm**	**Compared with 1 kPa**
kilopascal (kPa)	1 atm = 101.3 kPa	
millimeters of mercury (mm Hg)	1 atm = 760 mm Hg	1 kPa = 7.501 mm Hg
torr	1 atm = 760 torr	1 kPa = 7.501 torr
pounds per square inch (psi or lb/in^2)	1 atm = 14.7 psi	1 kPa = 0.145 psi
atmosphere (atm)		1 kPa = 0.009 869 atm

Units of pressure The SI unit of pressure is the pascal (Pa). It is named for Blaise Pascal, a French mathematician and philosopher. The pascal is derived from the SI unit of force, the newton (N), which is derived from three SI base units: the kilogram, the meter, and the second. One **pascal** is equal to a force of one newton per square meter: 1 Pa = 1 N/m^2. Many fields of science still use more traditional units of pressure. For example, engineers often report pressure as pounds per square inch (psi). The pressures measured by barometers and manometers can be reported in millimeters of mercury (mm Hg). There also is a unit called the torr, which is named to honor Torricelli. One torr is equal to one mm Hg.

At sea level, the average air pressure is 760 mm Hg when the temperature is 0°C. Air pressure often is reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mm Hg or 760 torr or 101.3 kilopascals (kPa). **Table 13-1** compares different units of pressure. Because the units 1 atm, 760 mm Hg, and 760 torr are defined units, they have as many significant figures as needed when used in calculations. Do the **problem-solving LAB** to see how the combined pressure of air and water affects divers.

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

problem-solving LAB

How are the depth of a dive and pressure related?

Making and Using Graphs For centuries, people have dived deep into the sea to collect items such as pearls. Nowadays, single-breath diving has become competitive. On January 18, 2000, Francisco "Pipin" Ferreras from Cuba set a new record with a dive of 162 meters. He was underwater for 3.2 minutes.

Analysis

Use the data in the table to make a graph of pressure versus depth.

Pressure Versus Depth	
Depth of dive (m)	**Pressure (atm)**
10	2.0
20	3.0
30	4.0
40	5.0
50	6.0
60	7.0

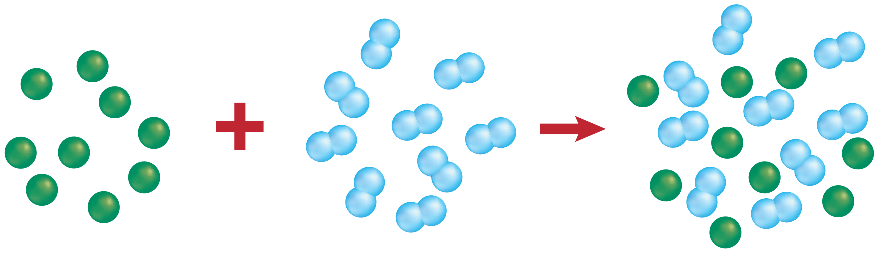
1. How are pressure and depth related?

2. What would the pressure be at the surface of the water? What does this value represent?

3. What was the pressure on Francisco at 162 m?

Thinking Critically

4. Using the equation for slope, calculate the slope of your graph. Then write an equation to express the relationship between pressure and depth. $\left(\text{slope} = \frac{\Delta y}{\Delta x} = \frac{y_2 - y_1}{x_2 - x_1}\right)$



1 mol He P_1 | 1 mol N_2 P_2 | 1 mol He + 1 mol N_2 P_{Total}

Figure 13-6

How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container? Refer to **Table C-1** in Appendix C for a key to atom color conventions.

Dalton's law of partial pressures When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of one mole of any gas is the same. Dalton's law of partial pressures can be summarized as

$$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$$

P_{total} represents the total pressure of a mixture of gases. P_1, P_2, and so on represent the partial pressures of each gas in the mixture. **Figure 13-6** shows what happens when one mole of helium and one mole of nitrogen are combined in a single closed container.

EXAMPLE PROBLEM 13-2

Finding the Partial Pressure of a Gas

A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2, if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1. Analyze the Problem

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$P_{N_2} = 0.12$ atm
$P_{CO_2} = 0.70$ atm
$P_{total} = 0.97$ atm

Unknown

$P_{O_2} = ?$ atm

2. Solve for the Unknown

Rearrange the equation to solve for the unknown value, P_{O_2}.

$P_{O_2} = P_{total} - P_{CO_2} - P_{N_2}$
$P_{O_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$
$P_{O_2} = 0.15$ atm

3. Evaluate the Answer

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure 0.97 atm. The answer has two significant figures to match the data.

Practice!

For more practice with partial pressure problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

4. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mm Hg and the partial pressure of helium is 439 mm Hg?

5. Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.

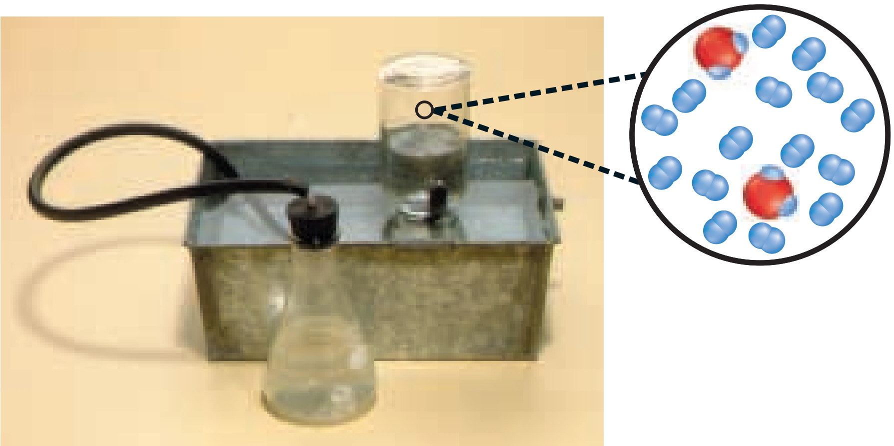
6. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.

Dalton's law of partial pressures can be used to determine the amount of gas produced by a reaction. The gas produced is bubbled into an inverted container of water, as shown in **Figure 13-7.** As the gas collects, it displaces the water. The gas collected in the container will be a mixture of hydrogen and water vapor. Therefore, the total pressure inside the container will be the sum of the partial pressures of hydrogen and water vapor.

The partial pressures of gases at the same temperature are related to their concentration. The partial pressure of water vapor has a fixed value at a given temperature. You can look up the value in a reference table. At 20°C, the partial pressure of water vapor is 2.3 kPa. You can calculate the partial pressure of hydrogen by subtracting the partial pressure of water vapor from the total pressure. If the total pressure of the hydrogen and water mixture is 95.0 kPa, what is the partial pressure of hydrogen at 20°C?

Figure 13-7

In the flask, sulfuric acid (H_2SO_4) reacts with zinc to produce hydrogen gas. The hydrogen is collected at 20°C.



As you will learn in Chapter 14, knowing the pressure, volume, and temperature of a gas allows you to calculate the number of moles of the gas. Temperature and volume can be measured during an experiment. Once the temperature is known, the partial pressure of water vapor is used to calculate the pressure of the gas. The known values for volume, temperature, and pressure are then used to find the number of moles.

Section 13.1 Assessment

7. What assumption of the kinetic-molecular theory explains why a gas can expand to fill a container?

8. How does the mass of a gas particle affect its rate of effusion?

9. Suppose two gases in a container have a total pressure of 1.20 atm. What is the pressure of gas B if the partial pressure of gas A is 0.75 atm?

10. Explain how changes in atmospheric pressure affect the height of the column of mercury in a barometer.

11. **Recognizing Cause and Effect** Explain why a tire or balloon expands when air is added.

12. **Thinking Critically** Explain why the container of water must be inverted when a gas is collected by displacement of water.

Forces of Attraction

Objectives

- **Describe** and **compare** intramolecular and intermolecular forces.
- **Distinguish** among intermolecular forces.

Vocabulary

dispersion force
dipole–dipole force
hydrogen bond

If all particles of matter at room temperature have the same average kinetic energy, why are some materials gases while others are liquids or solids? The answer lies with the attractive forces within and between particles. The attractive forces that hold particles together in ionic, covalent, and metallic bonds are called intramolecular forces. The prefix *intra-* means "within." For example, intramural sports are competitions among teams from within a single school. The term *molecular* can refer to atoms, ions, or molecules. **Table 13-2** summarizes what you learned about intramolecular forces in Chapters 8 and 9.

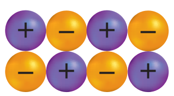
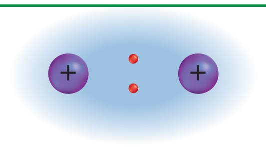
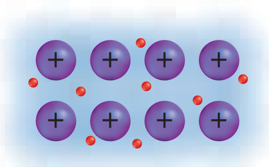
Intermolecular Forces

Intramolecular forces do not account for all attractions between particles. There are forces of attraction called intermolecular forces. The prefix *inter-* means "between" or "among." For example, an interview is a conversation between two people. Intermolecular forces can hold together identical particles, such as water molecules in a drop of water, or two different types of particles, such as carbon atoms in graphite and the cellulose particles in paper. The three intermolecular forces that will be discussed in this section are dispersion forces, dipole–dipole forces, and hydrogen bonds. Although some intermolecular forces are stronger than others, all intermolecular forces are weaker than intramolecular, or bonding, forces.

Dispersion forces Recall that oxygen molecules are nonpolar because electrons are evenly distributed between the equally electronegative oxygen atoms. Under the right conditions, however, oxygen molecules can be compressed into a liquid. For oxygen to be compressed, there must be some force of attraction between its molecules. The force of attraction between oxygen molecules is called a dispersion force. **Dispersion forces** are weak forces that result from temporary shifts in the density of electrons in electron clouds. Dispersion forces are sometimes called London forces after the German-American physicist who first described them, Fritz London.

Remember that the electrons in an electron cloud are in constant motion. When two nonpolar molecules are in close contact, especially when they collide, the

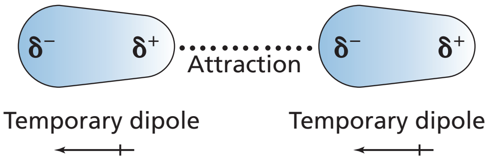
Table 13-2

Comparison of Intramolecular Forces			
Force	**Model**	**Basis of attraction**	**Example**
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H_2
Metallic		metal cations and mobile electrons	Fe

electron cloud of one molecule repels the electron cloud of the other molecule. The electron density around each nucleus is, for a moment, greater in one region of each cloud. Each molecule forms a temporary dipole.

Figure 13-8

What do the δ− and δ+ signs on a temporary dipole represent?



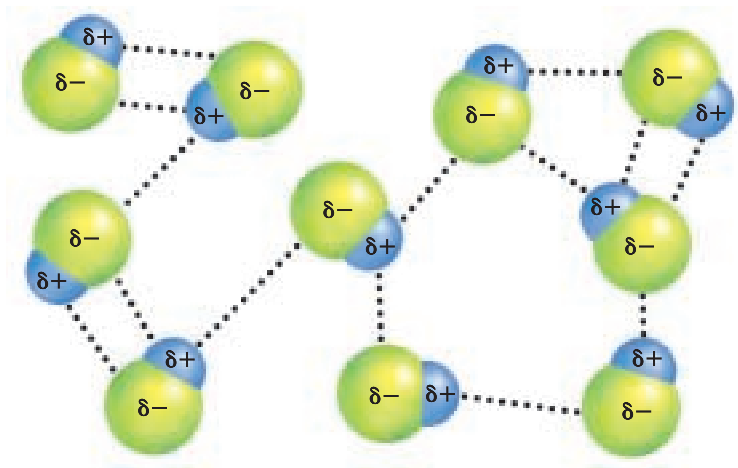
When temporary dipoles are close together, a weak dispersion force exists between oppositely charged regions of the dipoles, as shown in **Figure 13-8.**

Due to the temporary nature of the dipoles, dispersion forces are the weakest intermolecular force. Dispersion forces exist between all particles, but they play a significant role only when there are no stronger forces of attraction acting on particles. Dispersion forces are the dominant force of attraction between identical nonpolar molecules. These forces can have a noticeable effect as the number of electrons involved increases. For example, fluorine, chlorine, bromine, and iodine exist as diatomic molecules. Recall that the number of nonvalence electrons increases from fluorine to chlorine to bromine to iodine. Because the larger halogen molecules have more electrons, there can be a greater difference between the positive and negative regions of their temporary dipoles and, thus, stronger dispersion forces. This difference in dispersion forces explains why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature.

Dipole–dipole forces Polar molecules contain permanent dipoles; that is, some regions of a polar molecule are always partially negative and some regions of the molecule are always partially positive. Attractions between oppositely charged regions of polar molecules are called **dipole–dipole forces.** Neighboring polar molecules orient themselves so that oppositely charged regions line up. When hydrogen chloride gas molecules approach, the partially positive hydrogen atom in one molecule is attracted to the partially negative chlorine atom in another molecule. **Figure 13-9** shows multiple attractions among hydrogen chloride molecules.

Figure 13-9

The degree of polarity in a polar molecule depends on the relative electronegativity values of the elements in the molecule. What are the electronegativity values for hydrogen and chlorine?



Because the dipoles are permanent, you might expect dipole–dipole forces to be stronger than dispersion forces. This prediction will hold true as long as the molecules being compared have approximately the same mass.

Hydrogen bonds One special type of dipole–dipole attraction is called a hydrogen bond. A **hydrogen bond** is a dipole–dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair. For a hydrogen bond to form, hydrogen must be bonded to either a fluorine, oxygen, or nitrogen atom. These atoms are electronegative enough to cause a large partial positive charge on the hydrogen atom, yet small enough that their lone pairs of electrons can come close to hydrogen atoms. Consider, for example, a water molecule. In a water molecule, the hydrogen atoms have a large partial positive charge and the oxygen atom has a large partial negative charge. When water molecules approach, a hydrogen atom on one molecule is attracted to the oxygen atom on the other molecule, as shown in **Figure 13-10.**

Table 13-3

Properties of Three Molecular Compounds		
Compound	**Molar mass (g)**	**Boiling point (°C)**
Water (H_2O)	18.0	100
Methane (CH_4)	16.0	–164
Ammonia (NH_3)	17.0	–33.4

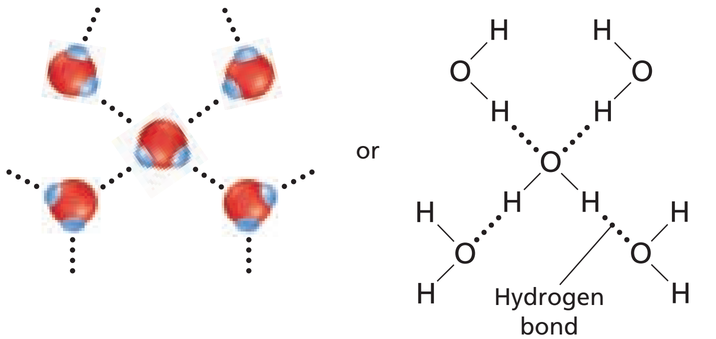


Figure 13-10

The hydrogen bonds between water molecules are stronger than typical dipole–dipole attractions because the bond between hydrogen and oxygen is highly polar.

Hydrogen bonds explain why water is a liquid at room temperature while compounds of comparable mass are gases. Look at the data in **Table 13-3**. The difference between methane and water is easy to explain. Because methane molecules are nonpolar, the only forces holding the molecules together are relatively weak dispersion forces. The difference between ammonia and water is not as obvious. Molecules of both compounds can form hydrogen bonds. Yet, ammonia is a gas at room temperature, which indicates that the attractive forces between ammonia molecules are not as strong. Because oxygen atoms are more electronegative than nitrogen atoms, the O–H bonds in water are more polar than the N–H bonds in ammonia. As a result, the hydrogen bonds between water molecules are stronger than those between ammonia molecules.

Section 13.2 Assessment

13. Why are dipole–dipole forces stronger than dispersion forces for molecules of comparable mass?

14. Which of the molecules listed below can form hydrogen bonds? For which of the molecules would dispersion forces be the only intermolecular force? Give reasons for your answers.

a. H_2
b. NH_3
c. HCl
d. HF

15. Predicting Make a prediction about the relative boiling points of the noble gases. Give a reason for your answer.

16. Thinking Critically In a methane molecule (CH_4), there are 4 single covalent bonds. In an octane molecule (C_8H_{18}), there are 25 single covalent bonds. How does the number of bonds affect the dispersion forces in samples of methane and octane? Which compound is a gas at room temperature? Which is a liquid?

Section 13.3 Liquids and Solids

Objectives

- **Apply** kinetic-molecular theory to the behavior of liquids and solids.
- **Relate** properties such as viscosity, surface tension, and capillary action to intermolecular forces.
- **Compare** the structures and properties of different types of solids.

Vocabulary

viscosity
surface tension
surfactant
crystalline solid
unit cell
amorphous solid

Although the kinetic-molecular theory was developed to explain the behavior of gases, the model can be applied to liquids and solids. When applying the kinetic-molecular theory to these states of matter, you must consider the forces of attraction between particles as well as their energy of motion.

Liquids

In Chapter 3, you learned that a liquid can take the shape of its container but that its volume is fixed. In other words, the particles in a liquid can flow to adjust to the shape of a container, but the liquid cannot expand to fill its container. Kinetic-molecular theory predicts the constant motion of the liquid particles. Individual liquid molecules do not have fixed positions in the liquid. However, forces of attraction between liquid particles limit their range of motion so that the particles remain closely packed in a fixed volume.

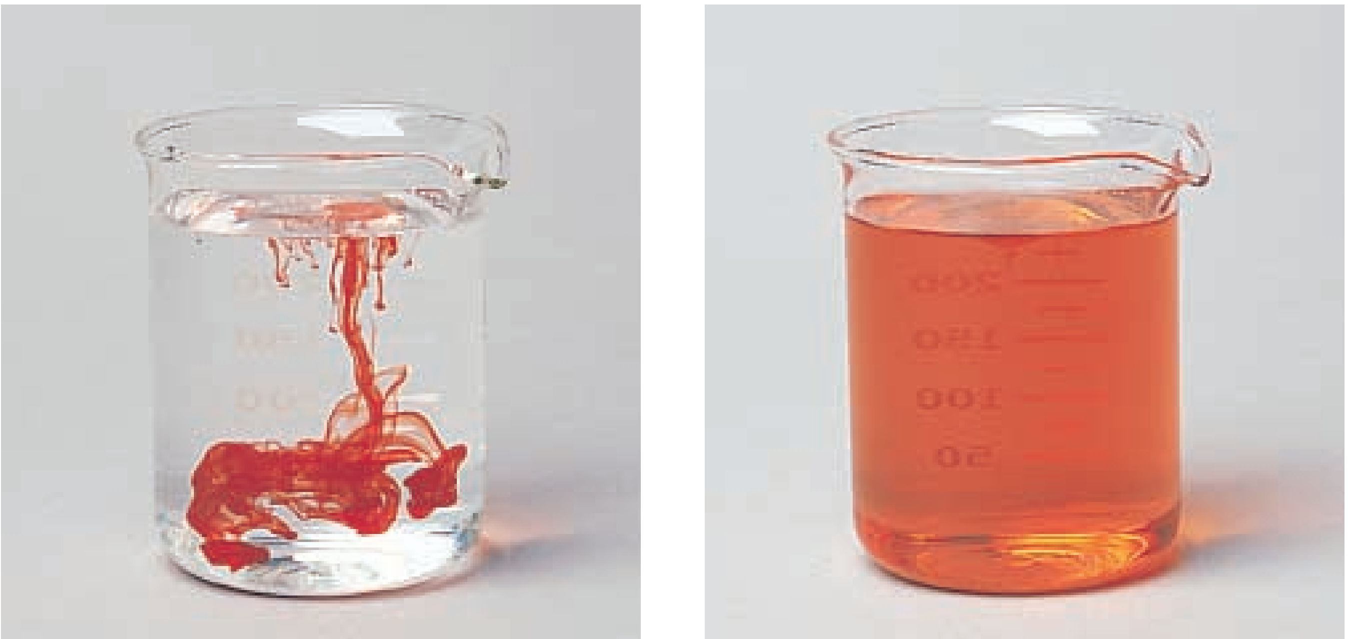
Density and compression At 25°C and one atmosphere of air pressure, liquids are much denser than gases. The density of a liquid is much greater than that of its vapor at the same conditions. For example, liquid water is about 1250 times denser than water vapor at 25°C and one atmosphere of pressure. Because they are at the same temperature, both gas and liquid particles have the same average kinetic energy. Thus, the higher density of liquids must be traced to the intermolecular forces that hold particles together.

Like gases, liquids can be compressed. But the change in volume for liquids is much smaller because liquid particles are already tightly packed together. An enormous amount of pressure must be applied to reduce the volume of a liquid by even a few percent.

Fluidity Fluidity is the ability to flow. Gases and liquids are classified as fluids because they can flow. A liquid can diffuse through another liquid as shown in **Figure 13-11.** A liquid diffuses more slowly than a gas at the same

Figure 13-11

How much time do you think it took for the food coloring to completely mix with the water?



temperature, however, because intermolecular attractions interfere with the flow. Thus, liquids are less fluid than gases. A comparison between water and natural gas can illustrate this difference. When there is a leak in a basement water pipe, the water remains in the basement unless the amount of water released exceeds the volume of the basement. Damage can be limited if the water supply is shut off quickly.

Natural gas, or methane, is the fuel burned in gas furnaces, gas hot-water heaters, and gas stoves. Gas that leaks from a gas pipe can diffuse throughout a house. Because natural gas is odorless, companies that supply the fuel include a compound with a distinct odor to warn customers of a leak. If the gas is effusing through a small hole in a pipe, the customer has time to shut off the gas supply, open windows to allow the gas to diffuse, and call the gas company to report the leak. If there is a break in a gas line, the customer must leave the house immediately because natural gas can explode when it comes in contact with an open flame or spark.

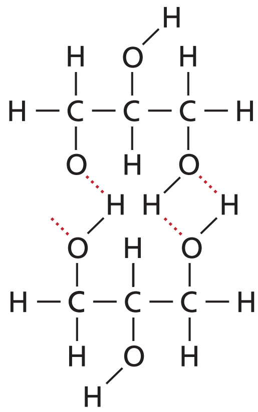


Figure 13-12

How many hydrogen bonds can a glycerol molecule form?

Viscosity Do you know the meaning of the phrase "slow as molasses"? Have you ever tried to get ketchup to flow out of a bottle? If so, you are already familiar with the concept of viscosity. **Viscosity** is a measure of the resistance of a liquid to flow. The particles in a liquid are close enough for attractive forces to slow their movement as they flow past one another. The viscosity of a liquid is determined by the type of intermolecular forces involved, the shape of the particles, and the temperature.

The stronger the attractive forces, the higher the viscosity. If you have used glycerol in the laboratory to help insert a glass tube into a rubber stopper, you know that glycerol is a viscous liquid. **Figure 13-12** uses structural formulas to show the hydrogen bonding that makes glycerol so viscous. Because it has three hydrogen atoms attached to oxygen atoms, a glycerol molecule can form three hydrogen bonds with other glycerol molecules.

The size and shape of particles affect viscosity. Recall that the overall kinetic energy of a particle is determined by its mass and velocity. Suppose the attractive forces between molecules in liquid A and liquid B are similar. If the molecules in liquid A are more massive than the molecules in liquid B, liquid A will have a greater viscosity. Liquid A's molecules will, on average, move more slowly than the molecules in liquid B. Molecules with long chains have a higher viscosity than shorter, more compact molecules, assuming the molecules exert the same type of attractive forces. Within the long chains, there is less distance between atoms on neighboring molecules and, thus, a greater chance for attractions between atoms. The long molecules in cooking oils and motor oils make these liquids thick and slow to pour.

Figure 13-13

The viscosity of motor oil increases in summer because additives in the oil change their shape as the temperature rises.

Viscosity and temperature Viscosity decreases with temperature. When you pour a tablespoon of cooking oil into a frying pan, the oil tends not to spread across the bottom of the pan until you heat the oil. With the increase in temperature, there is an increase in the average kinetic energy of the oil molecules. The added energy makes it easier for the molecules to overcome the intermolecular forces that keep the molecules from flowing.

In **Figure 13-13,** the motor oil that keeps the moving parts of an internal combustion engine lubricated is being replaced.

What could happen to an engine if there were no motor oil to lubricate its parts? Because temperature changes affect the viscosity of motor oil, people often used different motor oil blends in winter and summer. The motor oil used in winter was designed to keep flowing at low temperatures. The motor oil used in summer was more viscous so that it would maintain sufficient viscosity on extremely hot days or during long trips. Today, additives in motor oil can help adjust the viscosity so that the same oil blend can be used all year. Molecules in the additives are compact spheres with relatively low viscosity at cool temperatures. At high temperatures, the shape of the additive molecules changes to long strands. These strands get tangled with the oil molecules, which increases the viscosity of the oil.

Surface tension Intermolecular forces do not have an equal effect on all particles in a liquid, as shown in **Figure 13-14a.** Particles in the middle of the liquid can be attracted to particles above them, below them, and to either side. For particles at the surface of the liquid, there are no attractions from above to balance the attractions from below. Thus, there is a net attractive force pulling down on particles at the surface. The surface tends to have the smallest possible area and to act as though it is stretched tight like the head of a drum. For the surface area to increase, particles from the interior must move to the surface. It takes energy to overcome the attractions holding these particles in the interior. The energy required to increase the surface area of a liquid by a given amount is called **surface tension.** Surface tension is a measure of the inward pull by particles in the interior.

In general, the stronger the attractions between particles, the greater the surface tension. Water has a high surface tension because its molecules can form multiple hydrogen bonds. Drops of water are shaped like spheres because the surface area of a sphere is smaller than the surface area of any other shape of similar volume. Water's high surface tension is what allows the spider in **Figure 13-14b** to walk on the surface of a pond.

It is difficult to remove dirt from skin or clothing using only water. Because dirt particles cannot penetrate the surface of the water drops, the water cannot remove the dirt. What happened when you added a drop of detergent to the beaker in the **DISCOVERY LAB?** Soaps and detergents decrease the surface tension of water by disrupting the hydrogen bonds between water molecules. When the bonds are broken, the water spreads out. Compounds that lower the surface tension of water are called surface active agents or **surfactants.**

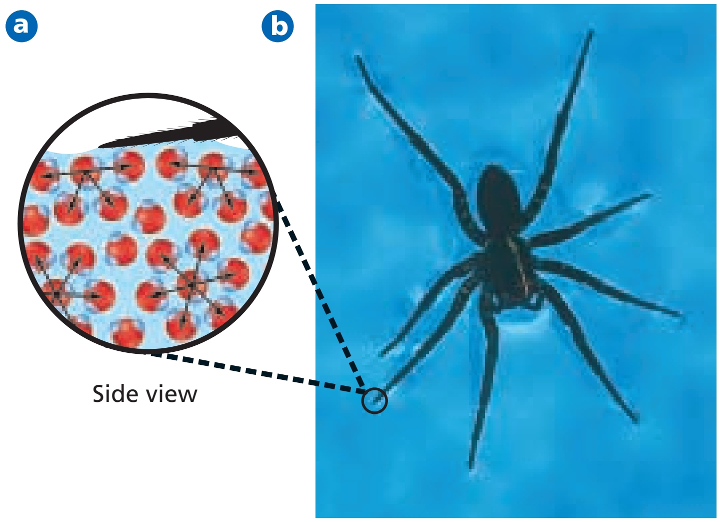


Figure 13-14

a Particles at the surface are drawn toward the interior of a liquid until attractive and repulsive forces are balanced.
b How does the spider's structure help it stay afloat on water?

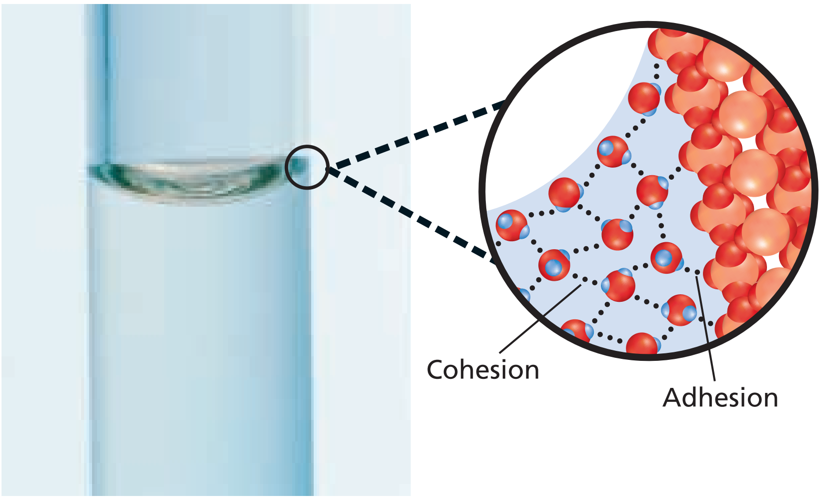


Figure 13-15

The surface of the water in a graduated cylinder is concave because water molecules are more strongly attracted to the silicon dioxide in glass than to other water molecules.

Capillary action When water is placed in a narrow container such as a graduated cylinder, you can see that the surface of the water is not straight. The surface forms a concave meniscus; that is, the surface dips in the center. **Figure 13-15** models what is happening to the water at the molecular level. There are two types of forces at work: cohesion and adhesion. Cohesion describes the force of attraction between identical molecules. Adhesion describes the force of attraction between molecules that are different. Because the adhesive forces between water molecules and the silicon dioxide in glass are greater than the cohesive forces between water molecules, the water rises along the inner walls of the cylinder.

If the cylinder is extremely narrow, a thin film of water will be drawn upward. Narrow tubes are called capillary tubes. This movement of a liquid such as water is called capillary action, or capillarity. Capillary action helps explain how paper towels can absorb large amounts of water. The water is drawn into the narrow spaces between the cellulose fibers in paper towels by capillary action. In addition, the water molecules form hydrogen bonds with cellulose molecules. These same factors account for the absorbent properties of disposable diapers. Water is drawn from the surface of the diaper to the interior by capillary action. The diaper can absorb about 200 times its mass in fluid.

Solids

According to the kinetic-molecular theory, a mole of solid particles has as much kinetic energy as a mole of liquid particles at the same temperature. By definition, the particles in a solid must be in constant motion. So why do solids have a definite shape and volume? For a substance to be a solid rather than a liquid at a given temperature, there must be strong attractive forces acting between particles in the solid. These forces limit the motion of the particles to vibrations around fixed locations in the solid. Thus, there is more order in a solid than in a liquid. Because of this order, solids are much less fluid than liquids and gases. In fact, solids are not classified as fluids.

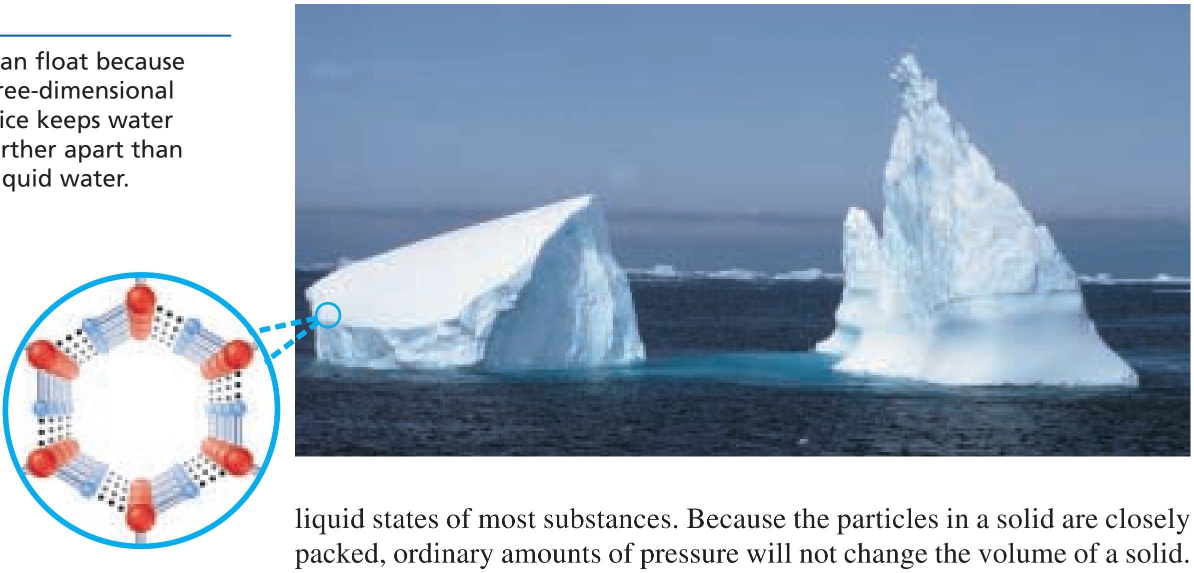
Density of solids In general, the particles in a solid are more closely packed than those in a liquid. Thus, most solids are more dense than most liquids. When the liquid and solid states of a substance coexist, the solid almost always sinks in the liquid. In **Figure 13-16,** solid cubes of benzene sink in liquid benzene because solid benzene is more dense than liquid benzene. There is about a 10% difference in density between the solid and

Figure 13-16

Ice cubes float in water. Benzene cubes sink in liquid benzene because solid benzene is more dense than liquid benzene.

Figure 13-17

An iceberg can float because the rigid, three-dimensional structure of ice keeps water molecules farther apart than they are in liquid water.



liquid states of most substances. Because the particles in a solid are closely packed, ordinary amounts of pressure will not change the volume of a solid.

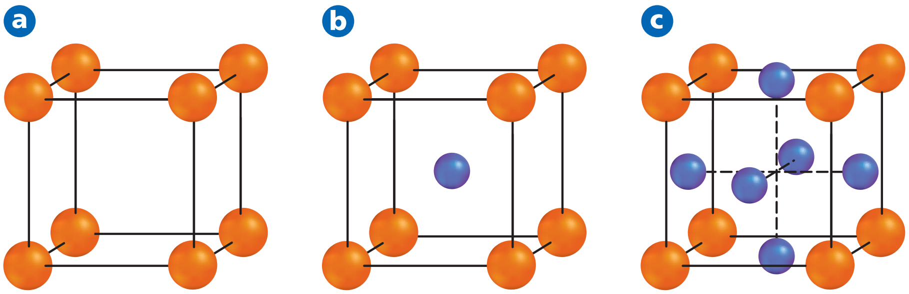
You cannot predict the relative densities of ice and liquid water based on benzene. Ice cubes and icebergs float because water is less dense as a solid than it is as a liquid. **Figure 13-17** shows the reason for the exception. The water molecules in ice are less closely packed together than in liquid water; that is, there is more space between the molecules in ice. As a result, there are more particles per unit volume in liquid water than in solid water.

Crystalline solids Although ice is unusual in its density, ice is typical of most solids in that its molecules are packed together in a predictable way. A **crystalline solid** is a solid whose atoms, ions, or molecules are arranged in an orderly, geometric, three-dimensional structure. The individual pieces of a crystalline solid are called crystals. In Chapter 8, you learned that the locations of ions in an ionic solid can be represented as points on a framework called a crystal lattice. A **unit cell** is the smallest arrangement of connected points that can be repeated in three directions to form the lattice. The relationship of a unit cell to a crystal lattice is similar to that of a formula unit to an ionic compound. Both the unit cell and the formula unit are small, representative parts of a much larger whole. **Figure 13-18** shows three ways that atoms or ions can be arranged in a cubic unit cell.

The shape of a crystalline solid is determined by the type of unit cell from which its lattice is built. **Figure 13-19** shows seven categories of crystals based on shape. Crystal shapes differ because the surfaces, or faces, of unit cells do not always meet at right angles and the edges of the faces vary in length. In **Figure 13-19,** the edges are labeled a, b, and c; the angles at which the faces meet are labeled α, β, and γ. Use the drawing as you model crystal shapes in the **miniLAB.**

Figure 13-18

These drawings show three of the ways particles are arranged in crystal lattices. Each sphere represents a particle. **a** Particles are arranged only at the corners of the cube. **b** There is a particle in the center of the cube. **c** There are particles in the center of each of the six cubic faces, but no particle in the center of the cube itself.



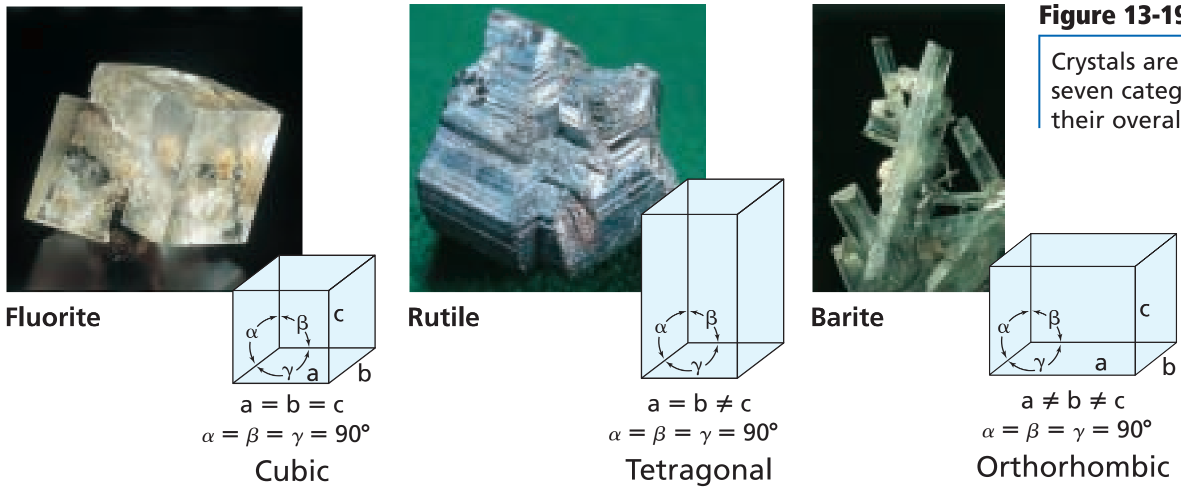
Simple cubic unit cell

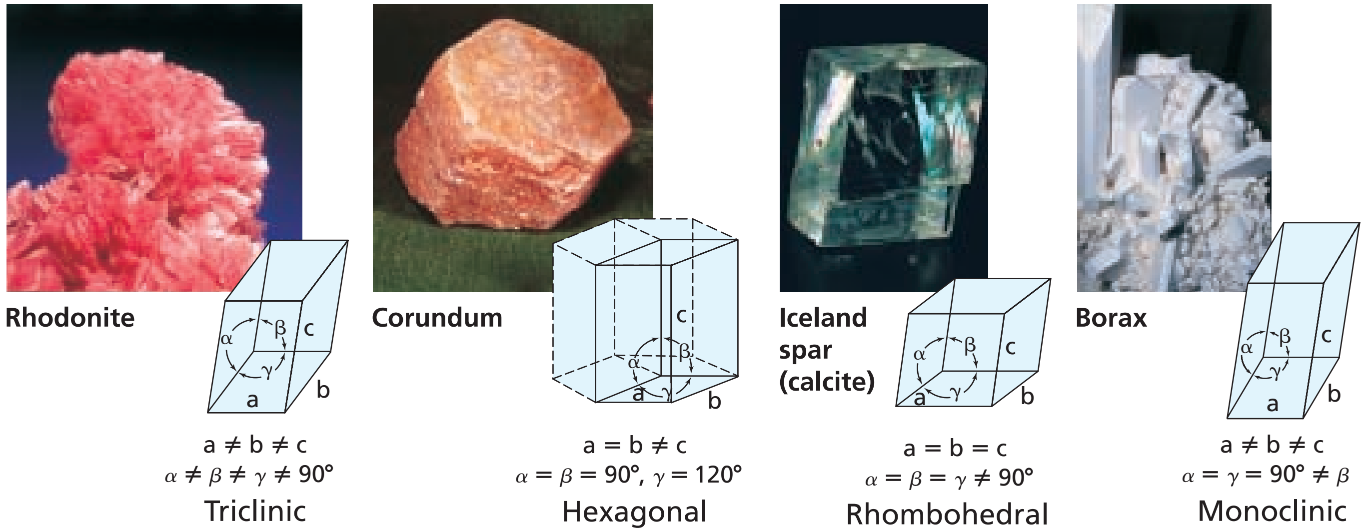
Body-centered cubic unit cell

Face-centered cubic unit cell

Figure 13-19

Crystals are classified into seven categories based on their overall shapes.





miniLAB

Crystal Unit Cell Models

Formulating Models You can make physical models that illustrate the structures of crystals.

Materials 12 plastic or paper soda straws, 22- or 26-gauge wire, scissors

Procedure

1. Cut four soda straws into thirds. Wire the pieces to make a cube. All angles are 90°.
2. To model a rhombohedral crystal, deform the cube from step 1 until no angles are 90°.
3. To model a hexagonal crystal, flatten the model from step 2 until it looks like a pie with six slices.
4. To model a tetragonal crystal, cut four straws in half. Cut four of the pieces in half again. Wire the eight shorter pieces to make four square ends. Use the longer pieces to connect the square ends.
5. To model the orthorhombic crystal, cut four straws in half. Cut 1/3 off four of the halves. Connect the four long, four medium, and four short pieces so that each side is a rectangle.
6. To model the monoclinic crystal, deform the model from step 5 along one axis. To model the triclinic crystal, deform the model from step 5 until it has no 90° angles.

Analysis

1. Which two models have three axes of equal length? How do these models differ?
2. Which model includes a square and rectangle?
3. Which models have three unequal axes?
4. Do you think crystals are perfect or do they have defects? Explain your answer.

Table 13-4

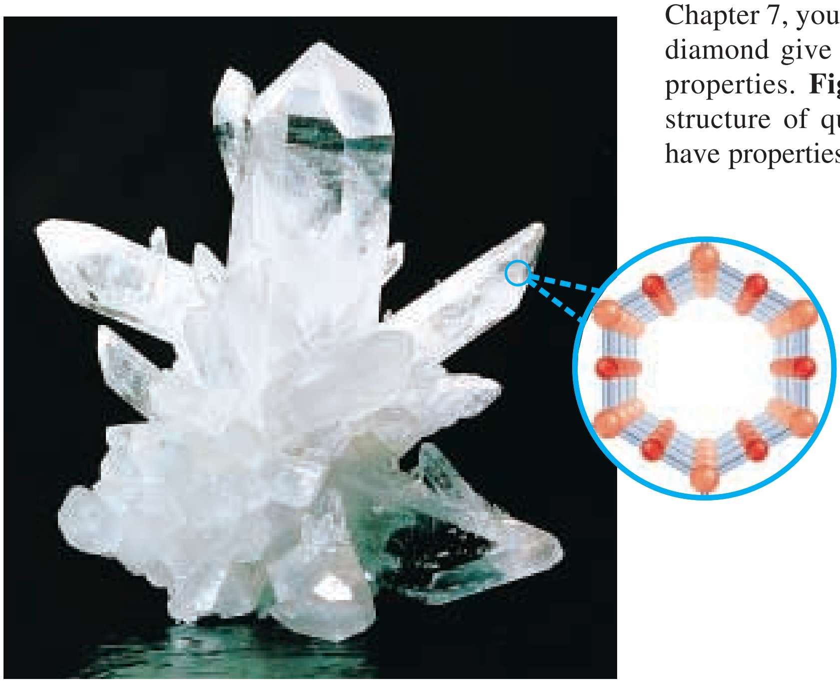
Types of Crystalline Solids			
Type	**Unit particles**	**Characteristics of solid phase**	**Examples**
atomic	atoms	soft to very soft; very low melting points; poor conductivity	group 8A elements
molecular	molecules	fairly soft; low to moderately high melting points; poor conductivity	I_2, H_2O, NH_3, CO_2, $C_{12}H_{22}O_{11}$ (table sugar)
covalent network	atoms connected by covalent bonds	very hard; very high melting points; often poor conductivity	diamond (C) and quartz (SiO_2)
ionic	ions	hard; brittle; high melting points; poor conductivity	NaCl, KBr, $CaCO_3$
metallic	atoms surrounded by mobile valence electrons	soft to hard; low to very high melting points; malleable and ductile; excellent conductivity	all metallic elements

Crystalline solids can be classified into five categories based on the types of particles they contain: atomic solids, molecular solids, covalent network solids, ionic solids, and metallic solids. **Table 13-4** summarizes the general characteristics of each category and provides examples. The only atomic solids are noble gases. Their properties reflect the weak dispersion forces between the atoms.

Molecular solids In molecular solids, the molecules are held together by dispersion forces, dipole–dipole forces, or hydrogen bonds. Most molecular compounds are not solids at room temperature. Even water, which can form strong hydrogen bonds, is a liquid at room temperature. Molecular compounds such as sugar are solids at room temperature because of their large molar masses. With larger molecules, many weak attractions can combine to hold the molecules together. Because they contain no ions, molecular solids are poor conductors of heat and electricity.

Figure 13-20

The most common kind of quartz has a hexagonal crystal structure.



Covalent network solids Atoms such as carbon and silicon, which can form multiple covalent bonds, are able to form covalent network solids. In Chapter 7, you learned how the structures of graphite and diamond give those solid allotropes of carbon different properties. **Figure 13-20** shows the covalent network structure of quartz. Based on its structure, will quartz have properties similar to diamond or graphite?

Ionic solids Remember that each ion in an ionic solid is surrounded by ions of opposite charge. The type of ions and the ratio of ions determine the structure of the lattice and the shape of the crystal. The network of attractions that extends throughout an ionic crystal gives these compounds their high melting points and hardness. Ionic crystals are strong, but brittle. When ionic crystals are struck, the cations and anions are shifted from their fixed positions. Repulsions between ions of like charge cause the crystal to shatter.

Figure 13-21

a The material a wire is made from, its thickness, and its length are variables that affect the flow of electrons through the wire. **b** Native Americans used the obsidian that forms when lava cools to make arrowheads and knives.

Metallic solids Recall from Chapter 8 that metallic solids consist of positive metal ions surrounded by a sea of mobile electrons. The strength of the metallic bonds between cations and electrons varies among metals and accounts for their wide range of physical properties. For example, tin melts at 232°C, but nickel melts at 1455°C. The mobile electrons make metals malleable—easily hammered into shapes—and ductile—easily drawn into wires. When force is applied to a metal, the electrons shift and thereby keep the metal ions bonded in their new positions. Read **Everyday Chemistry** at the end of the chapter to learn about shape-memory metals. Mobile electrons make metals good conductors of heat and electricity. Power lines carry electricity from power plants to homes and businesses and to the electric train shown in **Figure 13-21a.**

Amorphous solids Not all solids contain crystals. An **amorphous solid** is one in which the particles are not arranged in a regular, repeating pattern. The term amorphous is derived from a Greek word that means "without shape." An amorphous solid often forms when a molten material cools too quickly to allow enough time for crystals to form. **Figure 13-21b** shows an example of an amorphous solid.

Glass, rubber, and many plastics are amorphous solids. Recent studies have shown that glass may have some structure. When X-ray diffraction is used to study glass, there appears to be no pattern to the distribution of atoms. When neutrons are used instead, an orderly pattern of silicate units can be detected in some regions. Researchers hope to use this new information to control the structure of glass for optical applications and to produce glass that can conduct electricity.

Careers Using Chemistry

Materials Engineer

Would you like to change the atomic and molecular structure of substances to create products ranging from computer chips to skis? If so, consider a career as a materials engineer.

Materials engineers specialize in metals or ceramics. They create new materials to meet certain mechanical, electrical, or chemical standards. They also evaluate existing materials for new applications. Materials engineers work for the government and in the aviation, metal processing, electronics, automotive, and other industries in which materials play key roles.

Section 13.3 Assessment

17. Explain how hydrogen bonds affect the viscosity of a liquid. How does a change in temperature affect viscosity?

18. What effect does soap have on the surface tension of water?

19. How are a unit cell and a crystal lattice related?

20. Explain why solids are not classified as fluids.

21. What is the difference between a molecular solid and a covalent network solid?

22. Explain why most solids are denser than most liquids at the same temperature.

23. **Thinking Critically** Hypothesize why the surface of mercury in a thermometer is convex; that is, the surface is higher at the center.

Section 13.4 Phase Changes

Objectives

- **Explain** how the addition and removal of energy can cause a phase change.
- **Interpret** a phase diagram.

Vocabulary

melting point
vaporization
evaporation
vapor pressure
boiling point
sublimation
condensation
deposition
freezing point
phase diagram
triple point

Suppose you take a glass of ice water outside on a hot day. You set it down and rush inside to answer the phone. When you return much later, you find that the ice cubes have melted and there is less water in your glass. Can you use the kinetic-molecular theory to explain what happened to your drink?

Most substances can exist in three states depending on the temperature and pressure. A few substances, such as water, exist in all three states under ordinary conditions. States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture. Ice water is a heterogeneous mixture with two phases, solid ice and liquid water. When energy is added or removed from a system, one phase can change into another. As you read this section, use what you know about the kinetic-molecular theory to help explain the phase changes summarized in **Figure 13-22.**

Phase Changes That Require Energy

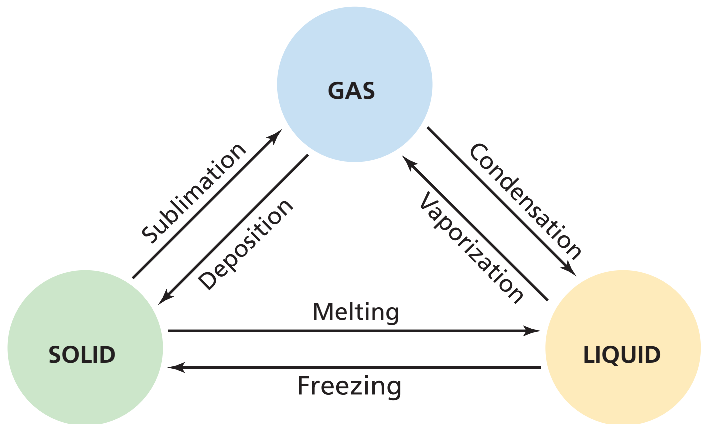
Because you are familiar with the phases of water—ice, liquid water, and water vapor—and have observed changes between those phases, we can use water as the primary example in the discussion of phase changes.

Melting What does happen to ice cubes in a glass of ice water? When ice cubes are placed in water, the water is at a higher temperature than the ice. Heat flows from the water to the ice. Heat is the transfer of energy from an object at a higher temperature to an object at a lower temperature. The energy absorbed by the ice is not used to raise the temperature of the ice. Instead, it disrupts the hydrogen bonds holding the water molecules together in the ice crystal. When molecules on the surface of the ice absorb enough energy to break the hydrogen bonds, they move apart and enter the liquid phase. As molecules are removed, the ice cube shrinks. The process continues until all of the ice melts. If a tray of ice cubes is left on a counter, where does the energy to melt the cubes come from?

The amount of energy required to melt one mole of a solid depends on the strength of the forces keeping the particles together in the solid. Because hydrogen bonds between water molecules are strong, a relatively large amount of energy is required. However, the energy required to melt ice is much less than the energy required to melt table salt because the ionic bonds in sodium chloride are much stronger than the hydrogen bonds in ice.

Figure 13-22

The six possible transitions between phases. What phase changes can occur between solids and liquids?



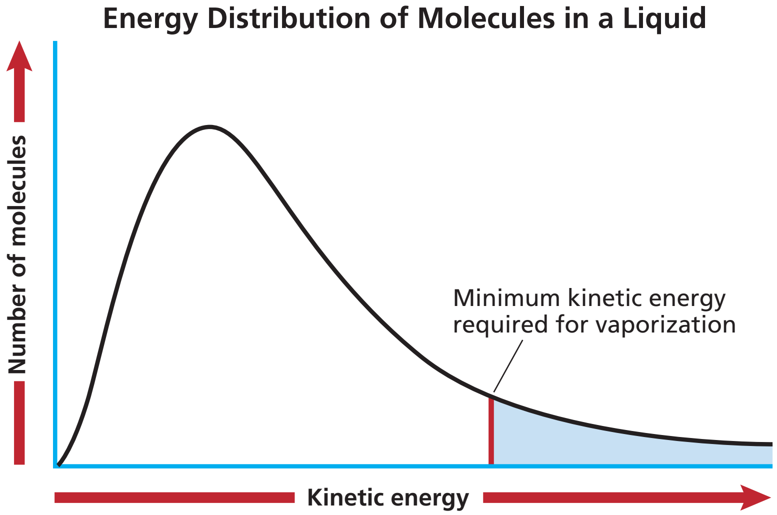


Figure 13-23

This graph shows a typical distribution of kinetic energies for a liquid at 25°C. The most probable kinetic energy lies at the peak of the curve. How would the curve look for the same liquid at 30°C?

The temperature at which the liquid phase and the solid phase of a given substance can coexist is a characteristic physical property of many solids. The **melting point** of a crystalline solid is the temperature at which the forces holding its crystal lattice together are broken and it becomes a liquid. It is difficult to specify an exact melting point for an amorphous solid because these solids tend to act like liquids when they are still in the solid state.

Vaporization While ice melts, the temperature of the ice–water mixture remains constant. Once all of the ice has melted, additional energy added to the system increases the kinetic energy of the liquid molecules. The temperature of the system begins to rise. In liquid water, some molecules will have more kinetic energy than other molecules. **Figure 13-23** shows how energy is distributed among the molecules in a liquid at 25°C. The shaded portion indicates those molecules that have the energy required to overcome the forces of attraction holding the molecules together in the liquid.

Particles that escape from the liquid enter the gas phase. For a substance that is ordinarily a liquid at room temperature, the gas phase is called a vapor. **Vaporization** is the process by which a liquid changes to a gas or vapor. If the input of energy is gradual, the molecules tend to escape from the surface of the liquid. Remember that molecules at the surface are attracted to fewer other molecules than are molecules in the interior. When vaporization occurs only at the surface of a liquid, the process is called **evaporation.** Even at cold temperatures, some water molecules have enough energy to evaporate. As the temperature rises, more and more molecules achieve the minimum energy required to escape from the liquid.

Evaporation, which requires energy, is the method by which your body controls its temperature. When you sit outside on a hot day or when you exercise, your body releases an aqueous solution called sweat from glands in your skin. Water molecules in sweat can absorb heat energy from your skin and evaporate. Excess heat is carried from all parts of your body to your skin by your blood. Evaporation of water also explains why a swim in cool ocean water is so refreshing. Not only is heat transferred from you to the cooler water while you are in the water, but water molecules left on your skin continue to cool you by evaporation once you come out of the water. The salts that remain when sea water evaporates often leave a white residue on your skin. To compare the rates of evaporation for different liquids, do the **CHEMLAB** at the end of the chapter.

Figure 13-24

Evaporation occurs in both open and closed containers.
a In an open container, water molecules that evaporate can escape from the container.
b Water vapor collects above the liquid in a closed container.

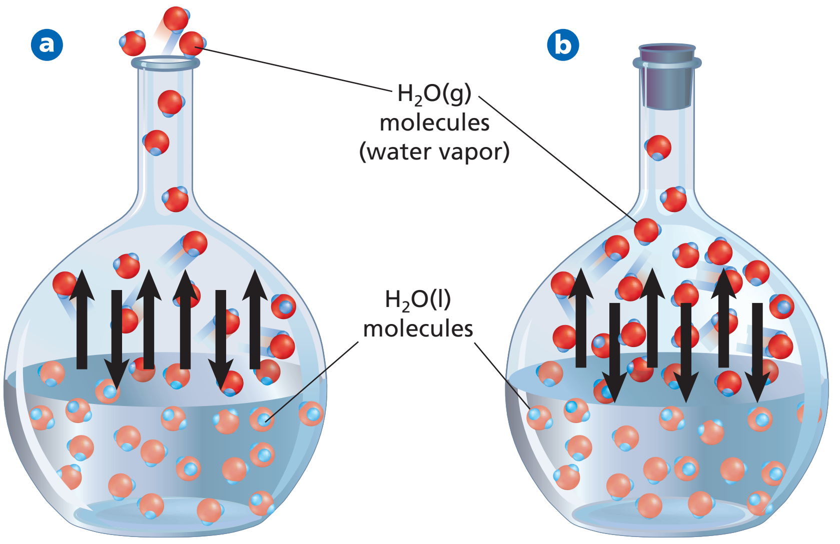


Figure 13-24 compares evaporation in an open container with evaporation in a closed container. If water is in an open container, all the molecules will eventually evaporate. The time it takes for them to evaporate depends on the amount of water and the available energy. How does temperature affect the rate of evaporation? In a partially filled, closed container, the situation is different. Water vapor collects above the liquid and exerts pressure on the surface of the liquid. The pressure exerted by a vapor over a liquid is called **vapor pressure.** How would a rise in temperature affect vapor pressure?

The temperature at which the vapor pressure of a liquid equals the external or atmospheric pressure is called the **boiling point.** Use **Figure 13-25** to compare what happens to a liquid at temperatures below its boiling point with what happens to a liquid at its boiling point. At the boiling point, molecules throughout the liquid have enough energy to vaporize. Bubbles of vapor collect below the surface of the liquid and rise to the surface. If a container has smooth walls and there are no dust particles to provide a site for bubble formation, a liquid can be heated to a temperature above its boiling point. When the liquid finally boils, the eruption of bubbles can cause the liquid to spatter. This problem can be avoided if a stone or ceramic chip is added to the liquid.

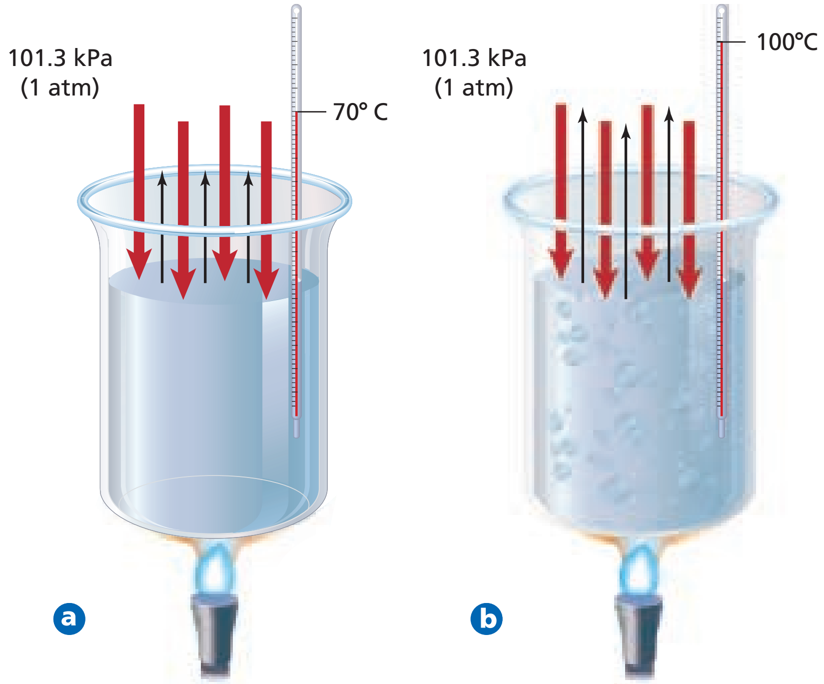


Figure 13-25

a As temperature increases, water molecules gain kinetic energy. Vapor pressure increases (black arrows) but is less than atmospheric pressure (red arrows). **b** A liquid has reached its boiling point when its vapor pressure is equal to atmospheric pressure. At sea level, the boiling point of water is 100°C.

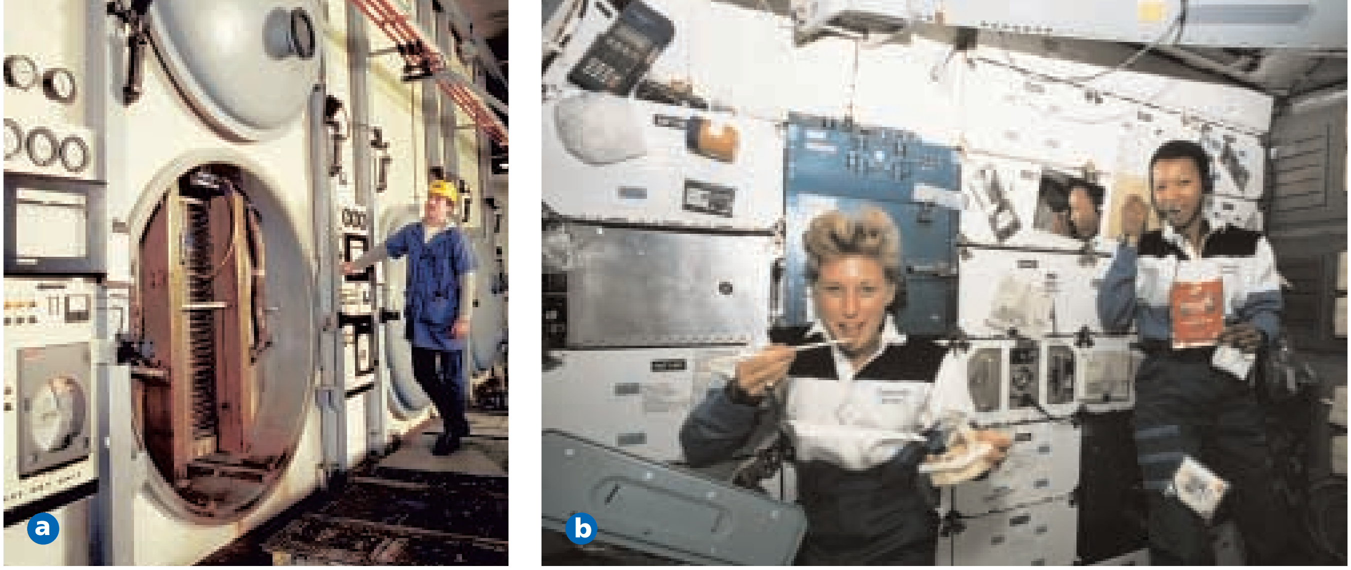


Figure 13-26

a Less flavor is lost when water is removed by freeze drying than when water is removed by evaporation. **b** These astronauts mix their freeze-dried food with water before eating.

Sublimation Many substances have the ability to change directly from the solid phase to the gas phase. **Sublimation** is the process by which a solid changes directly to a gas without first becoming a liquid. Solid iodine and solid carbon dioxide (dry ice) sublime at room temperature. The special effect shown in the chapter opener is caused when dry ice sublimes and cools water vapor in the air. Dry ice keeps objects that could be damaged by melting water cold during shipping. Moth balls, which contain the compounds napthalene or *p*-dichlorobenzene, sublime. So do solid air fresheners.

If ice cubes are left in a freezer for a long time, they shrink because the ice sublimes. At extremely low pressures, ice sublimes in a much shorter time period. This property of ice is used in a process called freeze drying. **Figure 13-26a** shows equipment used to produce freeze-dried food. Fresh food is frozen and placed in a container that is attached to a vacuum pump. As the pressure in the container is reduced, the ice sublimes and is removed from the container. The mass of the food is greatly reduced when water is removed. Backpackers often use freeze-dried foods because they want to carry as light a load as possible on a long hike. The astronauts shown in **Figure 13-26b** also have concerns about the mass of their cargo. More importantly, freeze-dried foods contain no water to support the growth of bacteria. They can be stored for a long time without refrigeration.

Phase Changes That Release Energy

Have you ever awakened on a chilly morning to see dew on the grass or frost on your windows? When you set a glass of ice water on a picnic table, do you notice beads of water on the outside of the glass? These events are examples of phase changes that release energy into the surroundings.

Condensation When a water vapor molecule loses energy, its velocity is reduced. The vapor molecule is more likely to interact and form a hydrogen bond when it collides with another water molecule. The formation of hydrogen bonds signals the change from the vapor phase to the liquid phase. Because liquid molecules are more dense than vapor molecules, the process by which a gas or a vapor becomes a liquid is called **condensation.** Condensation is the reverse of vaporization. When hydrogen bonds form in liquid water, energy is released.

There are different causes for the condensation of water vapor. All involve a transfer of energy. The vapor molecules can come in contact with a cold surface such as the outside of a glass containing ice water. Heat from the vapor

Figure 13-27

As water vapor cools, it can condense on a car or it can be deposited as six-sided snow crystals.

molecules is transferred to the glass as the water vapor condenses. The water vapor that condenses on blades of grass or the car shown in **Figure 13-27** forms liquid droplets called dew. When a layer of air near the ground cools, water vapor in the air condenses and produces fog. Dew and fog evaporate when exposed to sunlight. Clouds form when layers of air high above the surface of Earth cool. Clouds are made entirely of water droplets. When the drops grow large enough, they fall to the ground as rain.

Deposition Some substances can change directly into a solid without first forming a liquid. When water vapor comes in contact with a cold window in winter, it forms a solid deposit on the window called frost. **Deposition** is the process by which a substance changes from a gas or vapor to a solid without first becoming a liquid. Deposition is the reverse of sublimation. The snowflakes shown in **Figure 13-27** form when water vapor high up in the atmosphere changes directly into solid ice crystals. Energy is released as the crystals form.

Freezing Suppose you place liquid water in an ice tray in a freezer. As heat is removed from the water, the molecules lose kinetic energy and their velocity decreases. The molecules are less likely to flow past one another. When enough energy has been removed, the hydrogen bonds between water molecules keep the molecules fixed, or frozen, into set positions. Freezing is the reverse of melting. The **freezing point** is the temperature at which a liquid is converted into a crystalline solid. How do the melting point and freezing point of a given substance compare?

Earth Science CONNECTION

Sometimes, too much rain falls in one location and too little rain falls in another. People have been trying to produce rain on demand for centuries. Because most clouds exist at temperatures below the freezing point of water, rain often begins when water vapor deposits on ice crystals. (When ice crystals approach the surface of Earth, they melt and fall as rain if the temperature of the air near the surface is above freezing.) Rainmakers focus on the crucial role played by ice crystals. For example, dry ice pellets can be dropped into a cloud. The cold dry ice cools the water vapor in the cloud and ice crystals form. Sometimes tiny silver iodide crystals are sprayed into a cloud to serve as artificial "ice pellets."

Phase Diagrams

There are two variables that combine to control the phase of a substance: temperature and pressure. These variables can have opposite effects on a substance. For example, a temperature increase causes more liquid to vaporize, but an increase in pressure causes more vapor to condense. A **phase diagram** is a graph of pressure versus temperature that shows in which phase a substance exists under different conditions of temperature and pressure.

Figure 13-28a shows the phase diagram for water. You can use this graph to predict what phase water will be in for any combination of temperature and pressure. Note that there are three regions representing the solid, liquid, and vapor phases of water and three curves that separate the regions from one another. At points that fall along the curves, two phases of water can coexist. At what point on the graph do liquid water and water vapor exist at

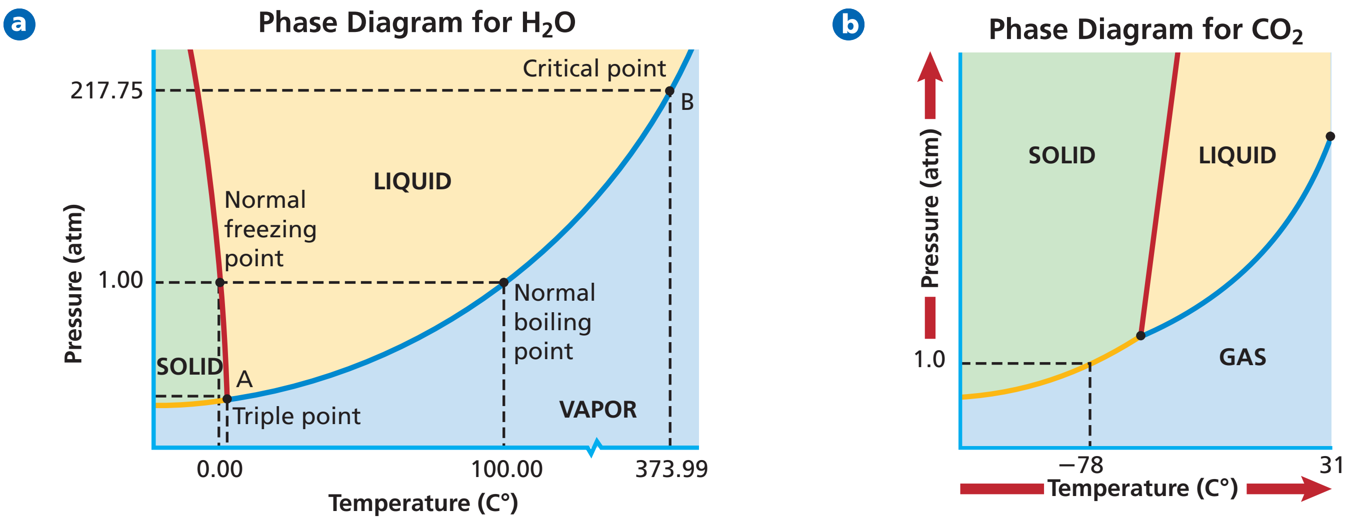


Figure 13-28

Compare the temperature and pressure values at the triple point for water and carbon dioxide.

1.00 atm and 100.00°C? The short, yellow curve shows the temperature and pressure conditions under which solid water and water vapor can coexist. The long, blue curve shows the temperature and pressure conditions under which liquid water and water vapor can coexist. The red curve shows the temperature and pressure conditions under which solid water and liquid water can coexist. What is the normal freezing point of water?

Point A on the phase diagram of water is the triple point for water. The **triple point** is the point on a phase diagram that represents the temperature and pressure at which three phases of a substance can coexist. All six phase changes can occur at the triple point: freezing and melting; evaporation and condensation; sublimation and deposition. Point B is called the critical point. This point indicates the critical pressure and critical temperature above which water cannot exist as a liquid. If water vapor is at the critical temperature, an increase in pressure will not change the vapor into a liquid.

The phase diagram for each substance is different because the normal boiling and freezing points of substances are different. However, each diagram will supply the same type of data for the phases, including a triple point. Of course, the range of temperatures chosen will vary to reflect the physical properties of the substance. Consider the phase diagram for carbon dioxide, which is shown in **Figure 13-28b.** If this diagram used the same temperature range chosen for water, there would be no solid region and an extrmely small liquid region. Find the triple point for carbon dioxide. Estimate the pressure and temperature conditions at the triple point.

Section 13.4 Assessment

24. What information does a phase diagram supply?

25. What is the major difference between the processes of melting and freezing?

26. Explain what the triple point and the critical point on a phase diagram represent.

27. **Comparing and Contrasting** Compare what happens to the energy, order, and spacing of particles when a solid other than ice changes to a liquid with what happens to the energy, order, and spacing of particles when a gas changes to a liquid.

28. **Thinking Critically** Aerosol cans contain compressed gases that, when released, help propel the contents out of the can. Why is it important to keep aerosol cans from overheating?

CHEMLAB 13 Small Scale

Comparing Rates of Evaporation

Several factors determine how fast a sample of liquid will evaporate. The volume of the sample is a key factor. A drop of water takes less time to evaporate than a liter of water. The amount of energy supplied to the sample is another factor. In this lab, you will investigate how the type of liquid and temperature affect the rate of evaporation.

Problem

How do intermolecular forces affect the evaporation rates of liquids?

Objectives

- **Measure** and **compare** the rates of evaporation for different liquids.
- **Classify** liquids based on their rates of evaporation.
- **Predict** which intermolecular forces exist between the particles of each liquid.

Materials

distilled water
ethanol
isopropyl alcohol
acetone
household ammonia
5 droppers
5 small plastic cups
grease pencil or masking tape and a marking pen
paper towel
square of waxed paper
stopwatch

Safety Precautions

- **Always wear safety goggles and a lab apron.**
- **Wear gloves because some of the liquids can dry out your skin.**
- **Avoid inhaling any of the vapors, especially ammonia.**
- **There should be no open flames in the lab; some of the liquids are flammable.**

Pre-Lab

1. Read the entire **CHEMLAB.** Prepare a data table similar to the one shown.
2. What is evaporation? Describe what happens at the molecular level during evaporation.
3. List the three possible intermolecular forces. Which force is the weakest? Which force is the strongest?
4. Look at the materials list for this lab. Consider the five liquids you will test. Predict which liquids will evaporate quickly and which will take longer to evaporate. Give reasons for your predictions.
5. To calculate an evaporation rate, you would divide the evaporation time by the quantity of liquid used. Explain why it is possible to use the evaporation times from this lab as evaporation rates.
6. Make sure you know how to use the stopwatch provided. Will you need to convert the reading on the stopwatch to seconds?

Evaporation Data

Liquid	Evaporation time (s)	Shape of liquid drop
distilled water		
ethanol		
warm ethanol		
isopropyl alcohol		
acetone		
household ammonia		

Procedure

1. Use a grease pencil or masking tape to label each of 5 small plastic cups. Use A for distilled water, B for ethanol, C for isopropyl alcohol, D for acetone, and E for household ammonia.
2. Place the plastic cups on a paper towel.

3. Use a dropper to collect about 1 mL of distilled water and place the water in the cup labeled A. Place the dropper on the paper towel directly in front of the cup. Repeat with the other liquids.

4. Place a square of waxed paper on your lab surface. Plan where on the waxed paper you will place each of the 5 drops that you will test. The drops must be as far apart as possible to avoid mixing.
5. Have your stopwatch ready. Collect some water in your water dropper and place a single drop on the waxed paper. Begin timing. Time how long it takes for the drop to completely evaporate. While you wait, make two drawings of the drop. One drawing should show the shape of the drop as viewed from above. The other drawing should be a side view at eye level. If the drop takes longer than 5 minutes to evaporate, record > 300 in your data table.
6. Repeat step 5 with the four other liquids.
7. Use the above procedure to design an experiment in which you can observe the effect of temperature on the rate of evaporation of ethanol. Your teacher will provide a sample of warm ethanol. Record your observations.

Cleanup and Disposal

1. Crumple up the waxed paper and place it in the container assigned by your teacher.
2. Place unused liquids in the containers specified by your teacher.
3. Wash out all droppers and test tubes except those used for distilled water.

Analyze and Conclude

1. **Classifying** Which liquids evaporated quickly? Which liquids were slow to evaporate?
2. **Drawing a Conclusion** Based on your data, in which liquid(s) are the attractive forces between molecules most likely to be dispersion forces?
3. **Interpreting Data** Make a generalization about the shape of a liquid drop and the evaporation rate of the liquid.
4. **Recognizing Cause and Effect** What is the relationship between surface tension and the shape of a liquid drop? What are the attractive forces that increase surface tension?
5. **Applying Concepts** The isopropyl alcohol you used is a mixture of isopropyl alcohol and water. Would pure isopropyl alcohol evaporate more quickly or more slowly compared to the alcohol and water mixture? Give a reason for your answer.
6. **Thinking Critically** Household ammonia is a mixture of ammonia and water. Based on the data you collected, is there more ammonia or more water in the mixture? Use what you learned about the relative strengths of the attractive forces in ammonia and water to support your conclusion.
7. **Drawing a Conclusion** How does the rate of evaporation of warm ethanol compare to ethanol at room temperature? Use kinetic-molecular theory to explain your observations.
8. **Error Analysis** How could you change the procedure to make it more precise?

Real-World Chemistry

1. The vapor phases of liquids such as acetone and alcohol are more flammable than their liquid phases. For flammable liquids, what is the relationship between evaporation rate and the likelihood that the liquid will burn?
2. Suggest why a person who has a higher than normal temperature might be given a rubdown with rubbing alcohol (70% isopropyl alcohol).
3. Table salt can be collected from salt water by evaporation. The water is placed in large shallow containers. What advantage do these shallow containers have over deep containers with the same overall volume?

Everyday Chemistry

Metals with a Memory

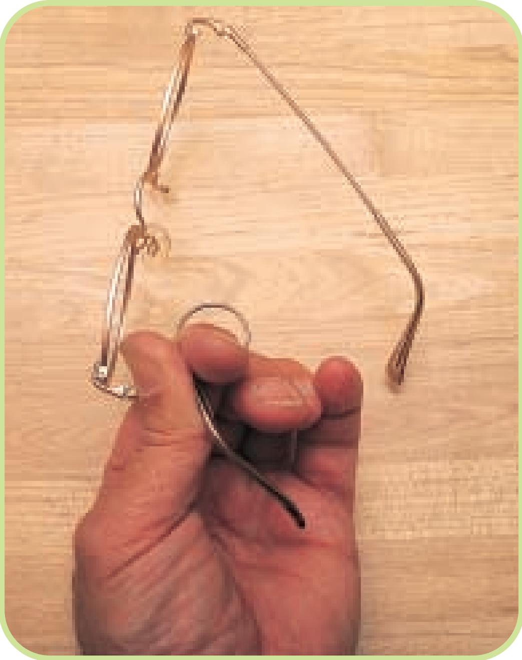
How can the eyeglass frames shown in the photo "remember" their original shape? How do braces move your teeth? These items may be made of a shape-memory alloy that has a remarkable property. It reverts to its previous shape when heated or when the stress that caused its shape is removed.

Two solid phases?

Melting is the transition of a material from a solid to a liquid. Transitions from one phase to another also can take place within a solid. A solid can have two phases if it has two possible crystal structures. It is the ability to undergo these changes in crystalline structure that gives shape-memory alloys their properties.

For example, an alloy that contains equal amounts of two metals may have phases called austentite and martensite. In the austentite phase, the metal ions are arranged in a rigid crystalline cubic structure. In the martensite phase, the metal ions are arranged in a more flexible structure. When the internal organization changes, the properties of the alloy change.

The austentite phase can be changed into the martensite phase if the alloy is cooled below the transition temperature under controlled conditions. Copper-aluminum-nickel, nickel-titanium, and copper-zinc-aluminum are the most common shape-memory alloys.



Nitinol

Nitinol is an alloy of nickel and titanium that has the austentite phase structure. Each nickel atom is at the center of a cube of titanium, and a titanium atom is at the center of each cube of nickel atoms. If nitinol is shaped into a straight wire, heated, and cooled past the transition temperature, it converts into the martensite phase. Now its structure allows the nitinol to be bent by an external stress. If the wire is heated, the stress is released and it reverts back to its initial shape.

Applications of nitinol

Nitinol is a corrosion resistant, light-weight material that generates a large force when it returns to its original shape. It also is safe to use inside the body, that is, nitinol is biocompatible. You may need fewer trips to the orthodontist because the nitinol wires in your braces need to be tightened and adjusted less frequently. Staples to repair human bones, devices that trap blood clots, and anchors used to reattach tendons to bone in an injured shoulder may be made of nitinol. Nitinol is used in a variety of military, safety, and robotics applications. Life-saving devices such as antiscalding valves that automatically shut off water flow and fire sprinklers that respond more quickly to heat rely on shape-memory alloys. So do vibration-control devices in buildings and bridges.

Testing Your Knowledge

1. **Inferring** Twisted nitinol-wire eyeglass frames unbend at room temperature. Is the transition temperature of the frames above or below room temperature?
2. **Applying** Design a simple lever that could be raised and lowered smoothly using nitinol wires.

CHAPTER 13 STUDY GUIDE

Summary

13.1 Gases

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Because of the space between gas particles, the density of gases is low and gases can be compressed.
- Because there are no significant forces of attraction between gas particles, gases can diffuse and effuse at rates determined by the mass of the particles.
- A barometer measures the pressure gas particles in Earth's atmosphere exert against Earth's surface.
- The total pressure of a gas mixture is the sum of the partial pressures of each gas in the mixture.

13.2 Forces of Attraction

- The intramolecular forces that hold together ionic, covalent, and metallic bonds are stronger than intermolecular forces.
- Dispersion forces are weak intermolecular forces between temporary dipoles of nonpolar molecules.
- Dipole–dipole forces occur between polar molecules. A hydrogen bond is a strong dipole–dipole force between molecules in which hydrogen atoms are bonded to highly electronegative atoms.

13.3 Liquids and Solids

- Particles are in motion in liquids and solids, but the range of motion is limited by intermolecular forces. Because liquids and solids are denser than gases, they are not easily compressed.
- In general, viscosity increases as the temperature decreases and as intermolecular forces increase.
- Surface tension results from an uneven distribution of attractive forces. A liquid displays capillarity when adhesive forces are stronger than cohesive forces.
- Except for amorphous solids, solids are more ordered than liquids. Crystalline solids can be classified by shape and composition.

13.4 Phase Changes

- Melting, vaporization, and sublimation are changes that require energy. Freezing, condensation, and deposition are changes that release energy. The temperature of a system remains constant during a phase change.
- Evaporation happens at the surface of a liquid. Boiling occurs when a liquid's vapor pressure is equal to atmospheric pressure.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

Key Equations and Relationships

Kinetic energy: $KE = 1/2mv^2$
(p. 386)

Graham's law of effusion: $\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$
(p. 387)

Dalton's law of partial pressures: $P_{\text{total}} = P_1 + P_2 + P_3 + \ldots P_n$
(p. 391)

Vocabulary

- amorphous solid (p. 403)
- atmosphere (p. 390)
- barometer (p. 389)
- boiling point (p. 406)
- condensation (p. 407)
- crystalline solid (p. 400)
- Dalton's law of partial pressures (p. 391)
- deposition (p. 408)
- diffusion (p. 387)
- dipole–dipole forces (p. 394)
- dispersion forces (p. 393)
- elastic collision (p. 386)
- evaporation (p. 405)
- freezing point (p. 408)
- Graham's law of effusion (p. 387)
- hydrogen bond (p. 395)
- kinetic-molecular theory (p. 385)
- melting point (p. 405)
- pascal (p. 390)
- phase diagram (p. 408)
- pressure (p. 388)
- sublimation (p. 407)
- surface tension (p. 398)
- surfactant (p. 398)
- temperature (p. 386)
- triple point (p. 409)
- unit cell (p. 400)
- vaporization (p. 405)
- vapor pressure (p. 406)
- viscosity (p. 397)

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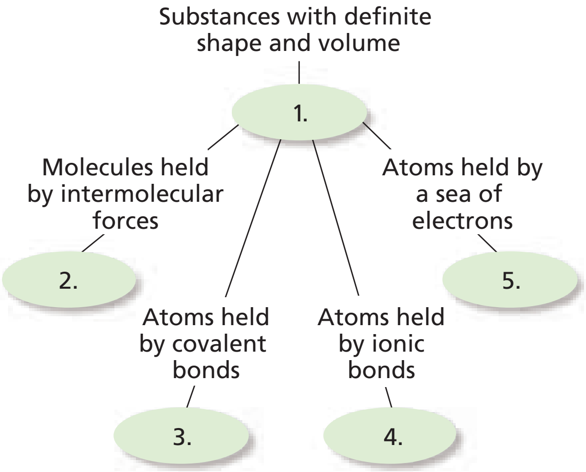
Concept Mapping

29. Complete the concept map using the following terms: covalent network solid, molecular solid, metallic solid, ionic solid, solid.



Mastering Concepts

30. What is an elastic collision? (13.1)

31. How does the kinetic energy of particles vary as a function of temperature? (13.1)

32. Use the kinetic-molecular theory to explain the compression and expansion of gases. (13.1)

33. Compare diffusion and effusion. Explain the relationship between the rates of these processes and the molar mass of a gas. (13.1)

34. What happens to the density of gas particles in a cylinder as the piston is raised? (13.1)

35. Explain why the baking instructions on a box of cake mix are different for high and low elevations. Would you expect to have a longer or shorter cooking time at a high elevation? (13.1)

36. Explain the difference between a temporary dipole and a permanent dipole. (13.2)

37. Why are dispersion forces weaker than dipole–dipole forces? (13.2)

38. Explain why hydrogen bonds are stronger than most dipole–dipole forces. (13.2)

39. Use relative differences in electronegativity to label the ends of the polar molecules listed as partially positive or partially negative. (13.2)

a. HF
b. NO
c. HBr
d. CO

40. Draw the structure of the dipole–dipole interaction between two molecules of carbon monoxide. (13.2)

41. Decide which of the substances listed can form hydrogen bonds. (13.2)

a. H_2O
b. HF
c. NaF
d. NO
e. H_2O_2
f. NH_3
g. H_2
h. CH_4

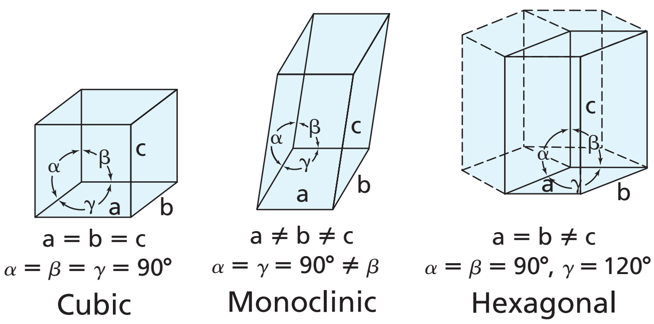
42. Hypothesize why long, nonpolar molecules would interact more strongly with one another than spherical nonpolar molecules of similar composition. (13.2)

43. What is surface tension and what conditions must exist for it to occur? (13.3)

44. Explain why the surface of water in a graduated cylinder is curved. (13.3)

45. Which liquid is more viscous at room temperature, water or molasses? Explain. (13.3)

46. Use these drawings to compare the cubic, monoclinic, and hexagonal crystal systems. (13.3)



47. What is the difference between a network solid and an ionic solid? (13.3)

48. Explain why most metals bend when struck but most ionic solids shatter. (13.3)

49. What is an amorphous solid? Under what conditions is such a solid likely to form? (13.3)

50. List the types of crystalline solids that are usually good conductors of heat and electricity. (13.3)

51. How does the strength of a liquid's intermolecular forces affect its viscosity? (13.3)

52. Explain why water has a higher surface tension than benzene, whose molecules are nonpolar. (13.3)

53. How does sublimation differs from deposition? (13.4)

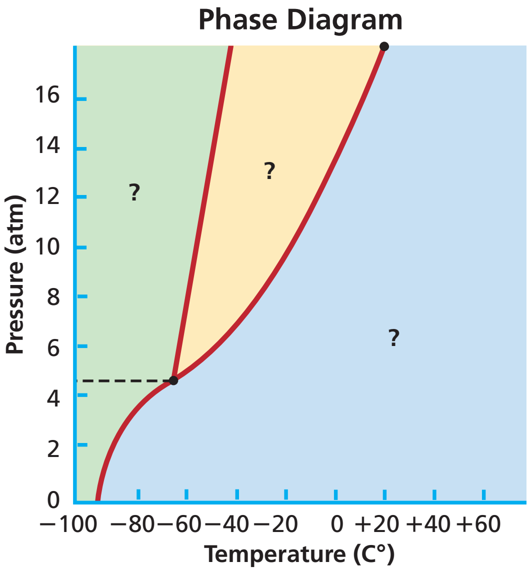
54. Compare boiling and evaporation. (13.4)

55. Define melting point. (13.4)

56. Explain the relationships among vapor pressure, atmospheric pressure, and boiling point. (13.4)

57. Explain why dew forms on cool mornings. (13.4)

58. Label the solid, liquid, and gas phases, triple point, and critical point on the phase diagram shown. (13.4)



59. Why does it take more energy to boil 10 g of liquid water than to melt an equivalent mass of ice? (13.4)

60. Why does a pile of snow slowly shrink even on days when the temperature never rises above the freezing point of water? (13.4)

61. Examine the phase diagram for carbon dioxide in **Figure 13-28.** Notice that at 1 atm pressure, the solid sublimes to a gas. What happens to the solid at much higher pressures? (13.4)

Mastering Problems

Graham's Law of Effusion (13.1)

62. What is the molar mass of a gas that takes three times longer to effuse than helium?

63. What is the ratio of effusion rates of krypton and neon at the same temperature and pressure?

64. Calculate the molar mass of a gas that diffuses three times faster than oxygen under similar conditions.

Dalton's Law of Partial Pressures (13.1)

65. What is the partial pressure of water vapor in an air sample when the total pressure is 1.00 atm, the partial pressure of nitrogen is 0.79 atm, the partial pressure of oxygen is 0.20 atm, and the partial pressure of all other gases in air is 0.0044 atm?

66. What is the total gas pressure in a sealed flask that contains oxygen at a partial pressure of 0.41 atm and water vapor at a partial pressure of 0.58 atm?

67. Find the partial pressure of oxygen in a sealed vessel that has a total pressure of 2.6 atm and also contains carbon dioxide at 1.3 atm and helium at 0.22 atm.

Converting Pressure Units (13.1)

68. What is the total pressure in atmospheres of a mixture of three gases with partial pressures of 12.0 kPa, 35.6 kPa, and 22.2 kPa?

69. The pressure atop the world's highest mountain, Mount Everest, is usually about 33.6 kPa. Convert the pressure to atmospheres. How does the pressure compare with the pressure at sea level?

70. The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?

71. At an ocean depth of 250 feet, the pressure is about 8.4 atm. Convert the pressure to mm Hg and kPa units.

Mixed Review

Sharpen your problem solving skills by answering the following.

72. Use the kinetic-molecular theory to explain why both gases and liquids are fluids.

73. Use intermolecular forces to explain why oxygen is a gas at room temperature and water is a liquid.

74. Use the kinetic-molecular theory to explain why gases are easier to compress than liquids or solids.

75. The density of mercury at 25°C and a pressure of 760 mm Hg is 13.5 g/mL; water at the same temperature and pressure has a density of 1.00 g/mL. Explain this difference in terms of intermolecular forces and the kinetic-molecular theory.

76. Two flasks of equal size are connected by a narrow tube that is closed in the middle with a stopcock. One flask has no gas particles; the other flask contains 0.1 mol of hydrogen gas at a pressure of 2.0 atm.

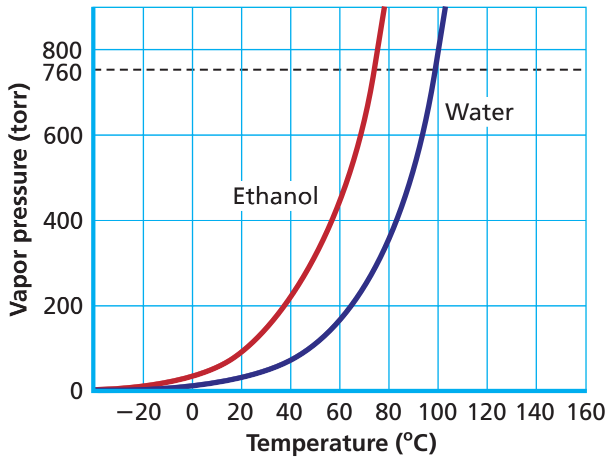
a. Describe what happens to the gas molecules after the stopcock is opened.

b. What will happen to the gas pressure after the stopcock is opened?

Thinking Critically

77. Interpreting Graphs Examine the graph below, which plots vapor pressure versus temperature for water and ethyl alcohol.

a. What is the boiling point of water at 1 atm?
b. What is the boiling point of ethyl alcohol at 1 atm?
c. Describe the relationship between temperature and vapor pressure for water and alcohol.
d. Estimate the temperature at which water will boil when the atmospheric pressure is 0.80 atm.



78. Applying Concepts A solid being heated stays at a constant temperature until it is completely melted. What happens to the heat energy put into the system during that time?

79. Comparing and Contrasting An air compressor uses energy to squeeze air particles together. When the air is released, it expands, allowing the energy to be used for purposes such as gently cleaning surfaces without using a more abrasive liquid or solid. Hydraulic systems essentially work the same way, but involve compression of liquid water rather than air. What do you think are some advantages and disadvantages of these two types of technology?

80. Hypothesizing What type of crystalline solid do you predict would best suit the following needs?

a. a material that can be melted and reformed at a low temperature
b. a material that can be drawn into long, thin wires
c. a material that conducts electricity when molten
d. an extremely hard material that is non-conductive

81. Communicating Which process is responsible for your being able to smell perfume from an open bottle that is located across the room from you, effusion or diffusion? Explain.

82. Inferring A laboratory demonstration involves pouring bromine vapors, which are a deep red color, into a flask of air and then tightly sealing the top of the flask. The bromine is observed to first sink to the bottom of the beaker. After several hours have passed, the red color is distributed equally throughout the flask.

a. Is bromine gas more or less dense than air?
b. Would liquid bromine diffuse more or less quickly than gaseous bromine after you pour it into another liquid?

Writing in Chemistry

83. Propane gas is a commonly used heating fuel for gas grills and homes. It is not packaged as a gas, however. It is liquefied and referred to as liquid propane or "LP gas." What advantages are there to storing and transporting propane as a liquid rather than a gas? Are there any disadvantages?

84. Find out what your birthstone is if you don't already know, and write a brief report about the chemistry of that gem. Find out its chemical composition, which category its unit cell is in, how hard and durable it is, and its approximate cost at the present time.

85. What would happen to life on Earth if ice were denser than liquid water? Would life be possible? Write an essay on this topic.

Cumulative Review

Refresh your understanding of the previous chapters by answering the following.

86. Identify the following as an element, compound, homogeneous mixture, or heterogeneous mixture. (Chapter 3)

a. air	**e.** ammonia
b. blood	**f.** mustard
c. antimony	**g.** water
d. brass	**h.** tin

87. Use the periodic table to separate these ten elements into five pairs of elements having similar properties. (Chapter 7)

S, Ne, Li, O, Mg, Ag, Sr, Kr, Cu, Na

88. You are given two clear, colorless aqueous solutions. You are told that one solution contains an ionic compound and one contains a covalent compound. How could you determine which is an ionic solution and which is a covalent solution? (Chapter 9)

89. Determine the number of atoms in 56.1 g Al. (Chapter 11)

STANDARDIZED TEST PRACTICE
CHAPTER 13

Use these questions and the test-taking tip to prepare for your standardized test.

1. Water has an extremely high boiling point compared to other compounds of similar molar mass because of ________ .

a. hydrogen bonding
b. adhesive forces
c. covalent bonding
d. dispersion forces

2. The ratio of effusion rates for nitric oxide (NO) and nitrogen tetroxide (N_2O_4) is ________ .

a. 0.326 **c.** 1.751
b. 0.571 **d.** 3.066

3. Which of the following is NOT an assumption of the kinetic-molecular theory?

a. Collisions between gas particles are elastic.
b. All the gas particles in a sample have the same velocity.
c. A gas particle is not significantly attracted or repelled by other gas particles.
d. All gases at a given temperature have the same average kinetic energy.

4. Which of the following statements does NOT describe what happens as a liquid boils?

a. The temperature of the system rises.
b. Energy is absorbed by the system.
c. The vapor pressure of the liquid is equal to atmospheric pressure.
d. The liquid is entering the gas phase.

5. The solid phase of a compound has a definite shape and volume because its particles ___________.

a. are not in constant motion.
b. are always packed more tightly than particles in the compound's liquid phase.
c. can only vibrate around fixed points.
d. are held together by strong intramolecular forces.

6. A sealed flask contains neon, argon, and krypton gas. If the total pressure in the flask is 3.782 atm, the partial pressure of Ne is 0.435 atm, and the partial pressure of Kr is 1.613 atm, what is the partial pressure of Ar?

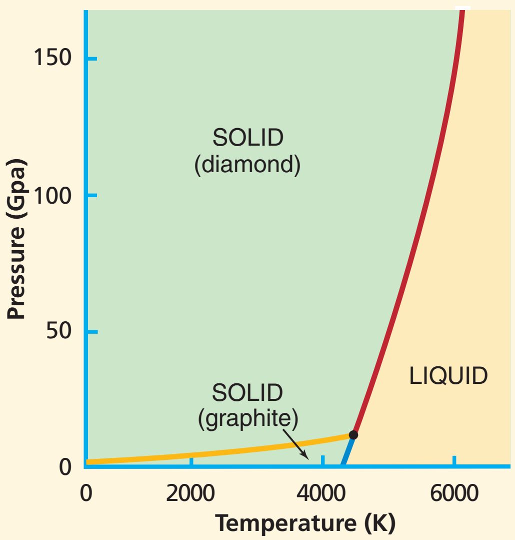
a. 2.048 torr **c.** 1556 torr
b. 1.734 torr **d.** 1318 torr

7. Which of the following does not affect the viscosity of a liquid?

a. intermolecular attractive forces
b. size and shape of molecules
c. temperature of the liquid
d. capillary action

Interpreting Graphs Use the graph to answer the following questions.

Phase Diagram of Carbon



8. Diamond is most likely to form at ________ .

a. temperatures > 5000 K and pressures < 100 GPa.
b. temperatures > 6000 K and pressures > 25 GPa.
c. temperatures < 4000 K and pressures > 25 GPa.
d. temperatures < 4500 K and pressures < 10 GPa.

9. Find the point on the graph at which carbon exists in three phases: solid graphite, solid diamond, and liquid carbon. The temperature and pressure at that point are ________ .

a. 4700 K and 15 GPa.
b. 3000 K and 10 GPa.
c. 5100 K and 50 GPa.
d. 14500 K and 5 GPa.

10. In what form or forms does carbon exist at 6000 K and 75 GPa?

a. diamond only
b. liquid carbon only
c. diamond and liquid carbon
d. liquid carbon and graphite

THE PRINCETON REVIEW **TEST-TAKING TIP**

Focus When you take a test, pay no attention to anyone other than the proctor. If students near you are talking, move to a different seat. If someone other than the proctor talks to you during a test, don't respond. Not only is talking a distraction, but the proctor may think that you are cheating. Don't take the chance. Focus on the test, and nothing else.